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-BUCKMASTER'S TEXT-BOOK-

INORGANIC CHEMISTRY

THE ELEMENTS
OF
INORGANIC CHEMISTRY.

BY
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OF THE SCIENCE AND ART DEPARTMENT, AND EXAMINER IN CHEMISTRY
AND PHYSICS TO THE ROYAL COLLEGE OF PRECEPTORS.

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PREFACE TO THE NINTH EDITION.

In preparing this Edition I have sought the best advice and assistance possible, so as to make the work suitable for instruction in Chemistry, according to the Syllabus prepared by the Science and Art Department. The most advanced ideas of Chemical Science have been introduced, the Constitutional and Graphic Formulæ adopted, and nothing omitted which is necessary for Elementary Instruction in Chemistry. The book will be found useful to students preparing for the matriculation of the London University and the Oxford and Cambridge Middle Class Examinations. In Schools where Chemistry is taught it is hoped that the work will be appreciated as a Text Book in which accurate and exact knowledge have not been sacrificed to popular explanation and experiment. The Syllabus and Exercises at the end will afford much assistance in directing the course of study.

supposed. Again, chalk, or calcic carbonate, is ma up of three elementary bodies,—a metal called calciu oxygen, and carbon. Common salt is composed two elementary bodies,—a metal called sodium, an a gas called chlorine. The great majority of compounds are formed by the union of three or four elementary bodies. The separation of these compound bodies into simpler forms is called *analysis*, and the formation of compound bodies by the union of simple ones is called *synthesis*.

The elementary bodies at present known are sixty-two in number; but many of them are so rare as not to require special attention. They are divided into *metals* and *non-metallic bodies*. The *metals* have great lustre: are good conductors of heat and electricity. These properties are never associated in a *non-metallic body*. This division is not based on any exact principles of science, but as a common means of description.

The elements are also divided into *chlorous*, or *negative*, and *basylous*, or *positive*: the former are evolved at the positive pole when [redacted] are decomposed

TABLE OF NON-METALLIC ELEMENTS.

Names.	Symbols.	Atomic Weights.	Specific Gravities.
Hydrogen	H	1	1
Chlorine	Cl	35·5	35·5
Bromine	Br	80	2·97
Iodine	I	127	4·948
Fluorine	F	19	...
Oxygen	O	16	16
Sulphur	S	32	2·07
Selenium	Se	79	4·28
Carbon	C	12	3·5
Silicon	Si	28·5	2·49
Boron	B	11	2·68
Nitrogen	N	14	14
Phosphorus	P	31	1·77

METALLIC ELEMENTS.

Names.	Symbols.	Atomic Weights.	Specific Gravities.
Potassium (*Kalium*) . .	K	39	0·865
Sodium (*Natrium*) . .	Na	23	0·972
Lithium	Li	7	·59
Barium	Ba	137	4 to 5
Strontium	Sr	87·5	2·5
Calcium	Ca	40	1·6
Magnesium	Mg	24	1·70
Aluminium	Al	27·5	2·60
Iron (*Ferrum*)	Fe	56	7·079
Chromium	Cr	52·5	5·9
Manganese	Mn	55	6·85
Zinc	Zn	65	7·1 to 7·86
Nickel	Ni	58·8	8·80
Cobalt	Co	58·8	8·54
Tin (*Stannum*). . . .	Sn	118	7·29
Antimony (*Stibium*) . .	Sb	122	6·70
Arsenic	As	75	5·88
Bismuth	Bi	208	9·82

METALLIC ELEMENTS (*continued*).

Names.	Symbols.	Atomic Weights.	Specific Gravities.
Cadmium	Cd	112	8·7
Copper (*Cuprum*) . .	Cu	63·5	8·96
Gold (*Aurum*) . . .	Au	196·7	19·5
Platinum	Pt	197·12	21·50
Lead (*Plumbum*) . . .	Pb	207	11·45
Mercury (*Hydrargyrum*) .	Hg	200	13·59
Silver (*Argentum*) . .	Ag	108	10·47

The rarer elements are—cæsium, cerium, didymium, glucinum, indium, iridium, lanthanium, molybdenum, niobium, osmium, palladium, rhodium, rubidium, ruthenium, tantalum, tellurium, thorium, thallium, titanium, tungsten or wolfram, uranium, vanadium, yttrium, zirconium.

Some of the elements have popular names: sulphur is called *brimstone;* mercury, *quicksilver*. Carbon when crystallized is a *diamond*, or *black-lead;* when prepared artificially it is known as *charcoal, coke,* tances. It also differs from cohesion, which binds like particles of matter together. The rain falls by gravitation, the particles of iron are held together by cohesion; but in iron rust (or oxide of iron) the oxygen is united to the iron by the force of chemical affinity. Chemical combination may be distinguished from mere mechanical mixture by the formation of compounds having properties for the most part different from those of the bodies which have entered into combination. And what is most remarkable, bodies of opposite chemical characters evince the strongest tendency to combination. Chemical affinity may be regarded as that force by which bodies unite, and form compounds having properties totally different from those of the constituents which have entered into combination. This combination is usually attended with an elevation of temperature. The elements are held together in combination with different degrees of force. Copper will displace mercury from its combination with chlorine, whilst the copper is in turn readily displaced by iron. From this it appears that chlorine has a greater affinity for copper than for mercury, and for iron than for copper.

The force of chemical affinity is very much influenced by heat. Some chemical compounds are entirely decomposed by heat; the particles seem to be removed beyond the sphere of chemical action by the repulsive force of heat. Heat destroys all organic substances, but this destruction gives rise to other compounds of a more permanent character. Light, also, modifies the force of affinity; if equal volumes of chlorine and hydrogen be mixed in the dark they do not unite, but under the influence of sunlight combination takes place with great violence. The particles of matter which combine are assumed to be in different electrical states, and this is regarded as the basis of all chemical action. Electrical currents are among some of the most powerful causes of chemical decomposition. As chemical combination can only take place when the particles or molecules are brought within the sphere of the chemical force, which acts at insensible distances, the force of cohesion must first be overcome; and this is done by dissolving the substance, usually in water, so that the particles may be brought into closer contact. Substances in a solid state rarely, if ever, combine, even if reduced to powder. When elements are liberated at the same time, in what is called their *nascent* condition, compounds are sometimes formed which cannot be effected under other circumstances. If nitrogen and hydrogen be mixed in the same vessel combination does not take place; but when these gases are set free at the same time, as in the case of manure heaps, ammonia is formed.

Weight.—By the term weight is meant the whole amount of matter a body contains, and this is generally expressed in grains, pounds, grammes, &c. *Specific* is taken as unity; and the specific gravities of gases are referred to that standard. Great skill is required in determining the specific gravity of gases or vapours, and nothing short of frequent practice can give accuracy in the results.

Atoms.—Chemists *assume* that the elementary bodies are built up of infinitely small particles, which they call atoms; they further *assume* that these atoms, with few exceptions, are all of the same size, but of different weights, and that this difference is expressed by the atomic weights of the elements. The exceptions are phosphorus and arsenic, whose atoms are believed to be half the usual size; and zinc, cadmium, and mercury, whose atoms are double the size. The theory of atoms was originated by the late Dr. Dalton, of Manchester, whose name it bears. *An atom is defined to be the least part of an elementary body which can enter into, or be expelled from, a compound.* The term is sometimes applied to those groups of atoms called compound radicals, such as ammonium, hydroxyl, &c.

Atomic Weights.—These are the atom weights of the elements, the atom weight of hydrogen being taken as unity. They are obtained in the case of elements which are gases, or volatile liquids or solids, by comparing the weight of a volume of the gas or vapour with the weight of the same volume of hydrogen under like conditions of temperature and pressure. Taking hydrogen as unity, then, *the densities of the elementary gases are identical with their atomic weights*, the elements named above being the only exceptions to this law. If atoms be of the same size, it will readily be understood that the same volume will hold the same number, and the difference in the weight of the volume will therefore express the difference in the weight of the individual atoms. For example: the weight of a litre* of hydrogen at 0° Centigrade and 760 millimetres (760 m.m.) barometric pressure equals ·0896 gramme. A litre of chlorine gas, under like conditions of temperature and pressure, weighs 35·5 × ·0896 gramme, and a litre of bromine vapour 80 × ·0896 gramme. *Assuming* that a litre holds the same number of atoms of hydrogen, chlorine, and bromine respectively, the individual atoms will stand related in weight to the numbers 1, 35·5, and 80.

In the case of fixed solids, such as metals which do not vaporize, the atomic weight is taken to be that weight of the solid which at any given temperature contains the same amount of heat as seven parts by weight of solid lithium at the same temperature. Thus it requires 56 parts of iron, 63·5 of copper, or 207 of lead, when heated to 100° C., to raise a litre of water through as many degrees of temperature as 7 parts of lithium would do. Therefore, 56, 63·5, and 207 are the atomic weights of iron, copper, and lead respectively.

uncombined state. All molecules are of the same size, for the law of Ampère, which most chemists now accept, states that "all gases and vapours contain the same number of molecules within the same volume." Many of the elementary molecules consist of pairs of atoms; the molecular weight, therefore, in these cases is twice the atomic weight.

The following are examples of elementary molecules:—

Monatomic molecules, containing one atom; mercury, cadmium, zinc.
Diatomic molecules, containing two atoms; hydrogen, chlorine, bromine, iodine, fluorine, oxygen, nitrogen, sulphur.
Triatomic molecules, containing three atoms; ozone (allotropic oxygen).
Tetratomic molecules, containing four atoms; phosphorus and arsenic.
Hexatomic molecules, containing six atoms; sulphur (at temperatures but little removed from its boiling point).

The molecules of compounds occupy the same volume as one molecule of hydrogen, no matter how many volumes the constituents may measure. Taking the molecule of hydrogen to measure *two volumes*, then—

1 vol. or atom of H, and 1 vol. or atom of Cl, combined form 2 vols. or 1 molecule of hydrochloric acid.
2 vols. or atoms of H, and 1 vol. or atom of O, combined form 2 vols. or 1 molecule of steam.
3 vols. or atoms of H, and 1 vol. or atom of N, combined form 2 vols. or 1 molecule of ammonia gas.
4 vols. or atoms of H, and 1 vol. or atom of C, combined form 2 vols. or 1 molecule of marsh gas.

Symbols, Formulæ.—The elements are represented by letters, which are called symbols; thus, S is the symbol for sulphur, P for phosphorus. When two or more elements have names commencing with the same letter, the single letter is assigned to the principal or non-metallic element, and a second letter in small type is attached to the initial letter for the others; thus, C is the symbol for carbon, Cl for chlorine, Cr for chromium, Ca for calcium. In some cases the initial letters of the Latin names are used; Ag (argentum) for silver, Fe (ferrum), iron, Hg (hydrargyrum), mercury.

The symbol of an element represents one atom or volume of that element; H represents one atom or one volume of hydrogen, Cl one atom or volume of chlorine. More atoms than one are indicated by placing a small figure at the right hand of the symbol; thus, H_2 means two atoms of hydrogen. A large figure placed in front of the symbol has also the same significance; as 2H, two atoms of hydrogen.

The molecules of compounds are represented by the juxtaposition of the symbols of their component elements; thus, HCl stands for a molecule of hydrochloric acid, OH_2 for one of water, NH_3 for one of ammonia gas: these groups of symbols are called *chemical formulæ.* Two or more molecules of a com-

Selenium, though not a metal, was thought to be one on its discovery, and its name has not been changed. Four of the non-metals have names terminating in "ine,"—chlorine, bromine, iodine, and fluorine; these are sometimes called halogens (salt-producers): three end in "on,"—carbon, boron, and silicon: other three terminate in "gen,"—oxygen, hydrogen, and nitrogen, the "gen" being derived from the Greek word γεννάω, to produce, and implying that the three substances are acid, water, and nitre producers respectively. The remaining names are either arbitrary, or are supposed to indicate some prominent property of the element.

Compounds.—The student of chemistry is often perplexed with the variety of names given by chemical authors to one and the same substance; as an example, if he consult a sufficient number of text-books he will find that the black compound produced by the union of oxygen with copper is variously called oxide of copper, protoxide of copper, cupric oxide, copper oxide; and if he attempt to purchase the article he will find that the shop name of it is black oxide of copper. This apparent confusion arises from the transitionary state of the science; but it is very desirable in the interest of science that this state of things should not continue. The name of a compound, on the one hand, should be sufficiently precise and expressive to enable us to identify it; and on the other hand, it should express as clearly as possible its composition and constitution,—the mode, namely, in which its atoms are arranged in its molecule. The latter condition unfortunately cannot always be fulfilled without the introduction of long and frequently not very euphonious words. In this work the systematic name will be followed by the shop or commercial name whenever it is thought desirable to introduce it; the latter name will be found in brackets.

The simplest form of a chemical compound is that in which two elements are united; these are called binary compounds. The name of the positive element, or some abbreviated form of it, or of its Latin name, is made to terminate in "ic," and is placed first; and the negative element, or its abbreviation, comes after it with the termination "ide." The following are examples of this class of compounds:—

A compound of—

Potassium	and	iodine	is called	potassic iodide.
Sodium	,,	chlorine	,,	sodic chloride.
Copper	,,	sulphur	,,	cupric sulphide.
Silver	,,	chlorine	,,	argentic chloride.
Lead	,,	oxygen	,,	plumbic oxide.
Calcium	,,	fluorine	,,	calcic fluoride.

The positive element frequently forms more than one compound with the same negative element; when this is the case, the compound which contains the largest proportion of the negative element has the

2 atoms of lead, and 1 atom of oxygen, form plumbous oxide.
1 atom ,, ,, ,, ,, ,, plumbic ,,
,, ,, ,, 2 atoms ,, ,, ,, peroxide.

When it is necessary to indicate the number of atoms of the positive and negative constituents respectively with great precision, the Greek numerals are used as prefixes; "mono" for one, "di" for two, "tri" for three, "tetra" or "tetr" for four, "penta" or "pent" for five, "hexa" or "hex" for six, "hepta" or "hept" for seven, and "octa" or "oct" for eight. Thus, a compound composed of two atoms of manganese and six atoms of chlorine is called—

*Di*manganic *hexa*chloride.

The student will do well to familiarize himself with the meanings of these Greek numerals, as they are very frequently used in the systematic names.

There are certain oxides which form acids when brought into contact with water; the usual nomenclature of binary compounds is not applied to these. They are called *anhydrides* (from α, negative, and υδωρ, water), or anhydrous acids; thus—

CO_2 is called carbonic anhydride.
SO_2 ,, sulphurous ,,
SO_3 ,, sulphuric ,,
N_2O_3 ,, nitrous ,,
N_2O_5 ,, nitric ,,
P_2O_5 ,, phosphoric ,,
SiO_2 ,, silicic ,,

These anhydrides may therefore be described as the residues of acids from which water has been abstracted.

Several of the hydrogen binary compounds still retain their old names, the systematic name being rarely used; thus—

Systematic Name.	Formula.	Common Name.
Hydric oxide	OH_2	water.
„ sulphide	SH_2	sulphuretted hydrogen.
„ chloride	HCl	hydrochloric acid.
„ bromide	HBr	hydrobromic acid.
„ iodide	HI	hydriodic acid.
„ fluoride	HF	hydrofluoric acid.
„ carbide	CH_4	marsh gas.
„ nitride	NH_3	ammonia gas.
„ phosphide	PH_3	phosphoretted hydrogen.
„ arsenide	AsH_3	arseniuretted hydrogen.

And there are a few oxides whose common and well-known names still keep their ground; thus—

Baric oxide, BaO, is more frequently called baryta.
Strontic „ SrO, „ strontia.
Calcic „ CaO, „ lime.
Magnesic „ MgO, „ magnesia.
Aluminic „ Al_2O_3, „ alumina.

Some of these oxides combine with water to form *hydrates;* thus—

or more atoms of hydrogen, which it can exchange for potassium or sodium when one of these metals is presented to it in the form of a hydrate." It is not always possible to exchange the whole of the hydrogen in this manner; and when this is the case the hydrogen which can be exchanged is called "replaceable hydrogen." An acid containing one such atom of replaceable hydrogen is said to be *monobasic;* when it contains two it is *dibasic;* three, *tribasic;* and four, *tetrabasic;* and we speak of its *basicity* being one, two, three, or four, as the case may be. When the basicity of an acid is greater than one, it is frequently spoken of as a *polybasic* acid.

The following examples will illustrate the mode in which the hydrogen of acids is replaced by sodium and potassium.*

HCl	$+$	OKH	$=$	KCl	$+$	OH_2.
Hydrochloric acid.		Potassic hydrate.		Potassic chloride.		Water.

In the above instance the potassium of the hydrate replaces the hydrogen of the acid.

NO_3H	$+$	$ONaH$	$=$	NO_3Na	$+$	OH_2.
Nitric acid.		Sodic hydrate.		Sodic nitrate.		Water.

Here the hydrogen of the nitric acid is exchanged for sodium.

SO_4H_2	$+$	$2ONaH$	$=$	SO_4Na_2	$+$	$2OH_2$.
Sulphuric acid.		Sodic hydrate.		Sodic sulphate.		Water.

And in this case the two atoms of hydrogen in the dibasic sulphuric acid are exchanged for two atoms of sodium.

---

* For an explanation of the meaning of the signs used in these equations, see p. 37.

There are two classes of acids: first, the *hydracids*, which contain no oxygen; they are formed by the combination of an element, such as one of the halogens, with hydrogen: thus we have hydrochloric acid, HCl; hydrobromic acid, HBr, &c. Their names are obtained by using the "ic" terminal to the chlorous element with the prefix "hydro" or "hydr." The other acids contain oxygen, and are hence distinguished by the name *oxy-acids;* they are named by adding the terminal "ic" to the name of the element with which the oxygen is united or to its abbreviation: thus we have nitric acid, NO_3H; chromic acid, CrO_4H_2, &c. When an element forms two oxy-acids the name of that which contains the smallest amount of oxygen has the terminal "ous:" thus we have sulphuric acid, SO_4H_2, and sulphurous acid, SO_3H_2. The prefix "per" is used to indicate an acid containing more oxygen than the one whose name terminates in "ic," and "hypo" is the prefix for those containing less oxygen than the one ending in "ous." Numerous examples of these acids will be found in the table of salts, p. 33 *et seq.*

There are some acids which contain sulphur in the place of oxygen; these are distinguished by having the prefix "sulpho" or "sulph" attached to them. Thus we have—

Sulpharsenious acid, AsS_3H_3.
Sulpharsenic acid, AsS_4H_3.

Bases.—A base is a compound which is converted into a salt by the action of an acid. In this action the elimination of water is the usual accompaniment.

Bases are of three classes:—

1. METALLIC OXIDES, such as sodic oxide, ONa_2,

potassic oxide, OK_2, calcic oxide, CaO, baric oxide BaO, cupric oxide, CuO, ferric oxide, Fe_2O_3.

2. Metallic hydrates, as sodic hydrate, $ONaH$, potassic hydrate, OKH calcic hydrate, $Ca(HO)_2$, ferric hydrate, $Fe_2(HO)_6$.

3. Certain compounds of hydrogen with nitrogen, phosphorus, arsenic, antimony, &c.: these are not of frequent occurrence; the nitrogen compound, ammonia, NH_3, being the most common.

The following examples will exhibit the mode whereby the bases are converted into salts: —

$$\underset{\text{Acid.}}{SO_4H_2} + \underset{\text{Base.}}{ONa_2} = \underset{\text{Salt.}}{SO_4Na_2} + \underset{\text{Water.}}{OH_2}.$$

$$\underset{\text{Acid.}}{NO_3H} + \underset{\text{Base.}}{OKH} = \underset{\text{Salt.}}{NO_3K} + \underset{\text{Water.}}{OH_2}.$$

$$\underset{\text{Acid.}}{SO_4H_2} + \underset{\text{Base.}}{Ca(HO)_2} = \underset{\text{Salt.}}{SO_4Ca} + \underset{\text{Water.}}{2OH_2}.$$

$$\underset{\text{Acid.}}{HCl} + \underset{\text{Base.}}{NH_3} = \underset{\text{Salt.}}{NH_4Cl}.$$

The bases of the first and second order are named according to the rules already given for binary compounds. The termination "ine" is used for the third order; as phosphine, PH_3, nitrine, NH_3; but the common names, phosphoretted hydrogen, ammonia, already given, are more frequently used to designate these bodies.

Salts.—It will already have been noticed from the foregoing examples that a salt is the product of the action of an acid upon a base; the hydracids yield in this manner *haloid* salts, the oxy-acids *oxy-salts*, and the sulpho-acids *sulpho-salts*.

Salts are also formed by the direct combination of metals with the halogens; thus, when sodium and other metals are burnt in chlorine gas, the chlorides of these metals are formed, and these are true salts.

The acids may be regarded as salts whose basylous element is hydrogen; in fact, each one of them is a body on the type of which a whole series of salts is framed: thus, hydrochloric acid (hydric chloride) is the type or representative of all chlorides, sulphuric acid (hydric sulphate) of all sulphates, nitric acid (hydric nitrate) of all nitrates.

The nomenclature of haloid salts follows the usual rule of binary compounds; thus we have potassic iodide, KI, calcic fluoride, CaF_2, ferric chloride, Fe_2Cl_6, &c.

The oxy-salts are named by adding the terminal "ate" to the name of the acid or its abbreviation when it ends in "ic," and the terminal "ite" when the name of the acid which forms it ends in "ous;" thus sulphuric acid forms sulphates, sulphurous acid sulphites.

Oxy-salts are said to be *acid*—or *normal salts*, according as the hydrogen of the acid is partially or completely replaced; thus, hydric-sodic sulphate, SO_4NaH, is an *acid* salt, because the hydrogen is only half replaced by sodium; but sodic sulphate, SO_4Na_2, is

The atoms of the elementary bodies have very different values in chemical exchange. This value is estimated in terms of an atom of hydrogen; thus, in the following equation one atom of sodium is found to take the place of one atom of hydrogen :—

$$\underset{\text{Hydrochloric acid.}}{HCl} + \underset{\text{Sodium.}}{Na} = \underset{\text{Sodic chloride.}}{NaCl} + \underset{\text{Hydrogen.}}{H}$$

In the above example one atom of sodium (23 parts by weight) replaces one atom of hydrogen (1 part by weight), and does an equal amount of combination work. Again, one atom of zinc in the following equation displaces two atoms of hydrogen, and therefore does twice the work of one atom of hydrogen :—

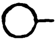
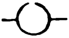
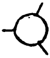
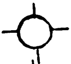
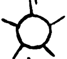
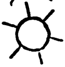
$$\underset{\text{Zinc.}}{Zn} + \underset{\text{Hydrochloric acid.}}{2HCl} = \underset{\text{Hydrogen.}}{H_2} + \underset{\text{Zincic chloride.}}{ZnCl_2}$$

When hydrogen and zinc are made to unite with chlorine, the hydrogen atom is found to be united to one atom of chlorine, and the zinc atom to two. It will be seen, therefore, that the atom-fixing and atom-replacing power of zinc is twice that of hydrogen.

Similarly an atom of gold can be substituted for three atoms of hydrogen, an atom of carbon for four, an atom of nitrogen for five, and one of sulphur for six respectively.

In order to account for this fact it is assumed that the atoms of different elements possess different powers, lines of force, or points of attachment; these have been named *bonds* by Dr. Frankland: thus the atoms of hydrogen and chlorine are single-bonded, whilst the atom of zinc has two such bonds, for it has twice the combining power of the hydrogen or chlorine atom.

To exhibit these facts *graphically* in the cases enumerated, the atoms may be represented by circles, and the bonds by short lines, thus :—

	Graphic representation of the atom with its bonds.
Hydrogen	
Zinc	
Gold	
Carbon	
Nitrogen	
Sulphur	

Dr. Odling was the first to introduce the use of dashes and Roman numerals to indicate the atomicity

A LIST OF THE MOST COMMON ELEMENTS,

Classified according to their Atomicities, with their Symbols and Atomic Weights (H = 1). The letters *a*, *b*, *c*, *d*, show a further subdivision into sections or families.

PERISSADS (Atomicity odd).

MONADS (1).			TRIADS (3).			PENTADS (5).		
a			*a*			*a*		
*Hydrogen	H	1	*Boron	B	11	*Nitrogen	N	14
b			*b*			*Phosphorus	P	31
†*Chlorine	Cl	35·5	Gold	Au	196·7	Arsenic	As	75
†*Bromine	Br	80				Antimony	Sb	122
†*Iodine	I	127				Bismuth	Bi	208
†*Fluorine	F	19						
c								
Potassium	K	39						
Sodium	Na	23						
Lithium	Li	7						
d								
Silver	Ag	108						

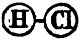
RTIADS (Atomicity even).

ETRADS (4).			HEXADS (6).		
a			*a*		
arbon .	C	12	†*Sulphur .	S	32
ilicon .	Si	28·5	†*Selenium .	Se	79
in .	Sn	118	*b*		
itanium .	Ti	50	Chromium .	Cr	52·5
b			Manganese .	Mn	55
luminium .	Al	27·5	Iron .	Fe	56
c			Cobalt .	Co	58·8
latinum .	Pt	197·12	Nickel .	Ni	58·8
alladium .	Pd	106·5	Uranium .	U	120
d					
ead .	Pb	207			

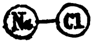
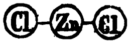
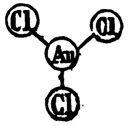
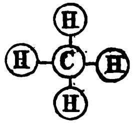
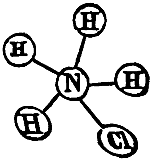
r their highest known atomicities. An artiad never becomes
, but the pentads are frequently triads or monads, and the

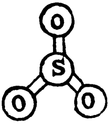
-metallic, the rest are metals. The elements marked † are

When two monad elements are in combination, the single bond of the one is occupied or fixed by the single bond of the other; thus the atoms of hydrogen and chlorine in combination are represented graphically as under :—



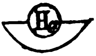
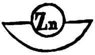
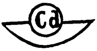
and this group of atoms constitutes the molecule of hydrochloric acid. A molecule may therefore be regarded as a group of atoms in which all the bonds are engaged. The following examples will illustrate this fact still further :—

	Symbolic formula.	Graphic formula.
Sodic chloride	$Na'Cl$	
Zincic chloride	$Zn''Cl_2$	
Auric chloride	$Au'''Cl_3$	
Marsh gas	$C^{IV}H_4$	
Ammonic chloride	$N^{V}H_4Cl$	

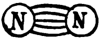
	Symbolic formula.	Graphic formula.
Sulphuric anhydride .	$S^{VI}O''_3$ . . .	

In the above cases the bonds of the elements are all occupied, and the molecules are complete ones.

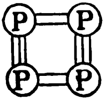
Elementary atoms cannot exist in a separate condition unless all their bonds are engaged. The two bonds of a diad atom are sometimes found to satisfy each other, and thus form a monatomic molecule; the following are well-known examples:—

Mercury .	. Hg'' . .	
Zinc .	. Zn'' . .	
Cadmium	. Cd'' . .	

Nitrogen . . . N^{v}_2 . .



Phosphorus . . P^{v}_4 . .

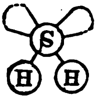


The elementary atoms are not always found to be united to the same proportion of another element: thus in sulphuretted hydrogen (SH_2) the sulphur has only two bonds occupied by the two monad atoms of hydrogen; in sulphurous anhydride (SO_2) the sulphur atom has four bonds saturated, for it is united to two diad atoms of oxygen; and in sulphuric anhydride (SO_3) all its bonds are engaged. Again, in ammonic chloride ($N^{v}H_4Cl$) all the five bonds of nitrogen are engaged; in ammonia gas (NH_3) only three of them; and in nitrous oxide (N_2O) one bond only of each nitrogen atom is fixed by one of the oxygen bonds.

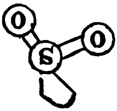
It will be noticed that in the above examples the bonds disappear in pairs, and it has been suggested by Dr. Frankland that this may be accounted for by assuming that the bonds not engaged are satisfying one another. How this is accomplished in the cases above cited may be represented graphically as follows:—

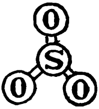
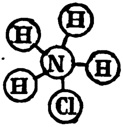
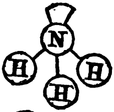
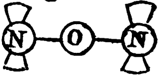
Symbolic. Graphic.

Sulphuretted hydrogen . . SH_2 . .



Sulphurous anhydride . . SO_2 . .



	Symbolic.	Graphic.
Sulphuric anhydride	SO_3	
Ammonic chloride	NH_4Cl	
Ammonia gas	NH_3	
Nitrous oxide	ON_2	

The highest number of bonds which an element is known to have is called its *absolute* atomicity, the number of bonds engaged in a compound is called respectively. As the bonds become latent in pairs, it will readily be understood that an artiad never can become a perissad, nor a perissad an artiad.

Compound Radicals.—The elements are simple radicals. There are certain groups of elementary bodies which are found in a great many chemical compounds, and which are a very important class of bodies; they perform all the functions of elementary bodies; they can be transferred from one compound to another in the same way that elements can. These bodies are called compound radicals. The following is a list of some of the most important of them, with their formulæ:

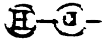
	Molecule.	Atom.	Abbreviated atomic formulæ.
Hydroxyl	$(HO)_2$	$(HO)'$	Ho
Hydrosulphyl	$(HS)_2$	$(HS)'$	Hs
Ammonium	$(NH_4)_2$	$(NH_4)'$	Am
Ammonoxyl	$(NH_4O)_2$	$(NH_4O)'$	Amo

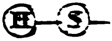
The metals also enter into the composition of these radicals, the metal being engaged by one bond only of the oxygen, which serves to link it to the other part of the molecule:—

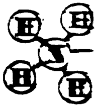
	Molecule.	Atom.	Abbreviated atomic formulæ.
Potassoxyl	$(KO)_2$	$(KO)'$	Ko'
Sodoxyl	$(NaO)_2$	$(NaO)'$	Nao'
Zincoxyl	$(ZnO_2)_2$	$(ZnO_2)''$	Zno'

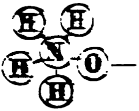
The compound radical atoms are groups of elements in which one or more bonds are disengaged; like elementary atoms they cannot exist in a separate condition, for when they are isolated, as some of them can be, they double themselves to form the molecule. They perform all the functions of elementary bodies: those which have only one bond unsatisfied are monad radicals, as sodoxyl; those with two bonds unsatisfied are diad radicals, as zincoxyl.

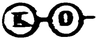
The following are the graphic formulæ of the above compound radicals:—

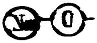
Hydroxyl . 

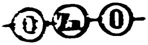
Hydrosulphyl . 

Ammonium . 

Ammonoxyl 

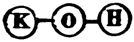
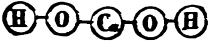
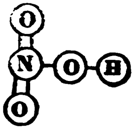
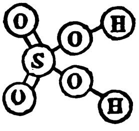
Potassoxyl . 

Sodoxyl . 

Zincoxyl . 

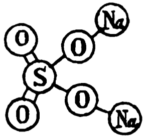
And the mode in which they enter into combination may be seen in a few selected formulæ of compounds: hydroxyl enters into the composition of all oxy-acids and hydrates.

Compounds containing hydroxyl:—

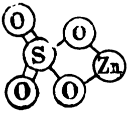
	Symbolic formulæ.	Graphic formulæ.
Potassic hydrate .	KHo . .	
Calcic hydrate . .	$CaHo_2$	
Nitric acid . . .	NO_2Ho .	
Sulphuric acid . .	SO_2Ho_2 .	

Compounds containing other radicals:—

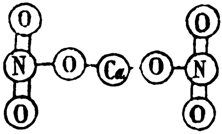
Sodic sulphate . . SO_2Nao_2



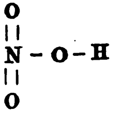
Zincic sulphate . . SO_2Zno'' .



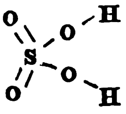
Calcic nitrate . . $\left\{\begin{matrix}NO_2\\ Cao''\\ NO_2\end{matrix}\right.$ . .



As the symbol of an element represents an atom of that element, it will readily be seen that the circles may be dispensed with in the graphic formulæ, the bond lines still being retained; thus nitric acid and sulphuric acid may be graphically represented as follows:—



$$\begin{matrix}O\\ ||\\ N-O-H\\ ||\\ O\end{matrix}$$



The student will therefore gradually abandon the use of the circle when it has fully served its purpose.

Notation.—There are at present three different forms of formulæ in use for the purpose of representing the atomic constitution of molecules, or the mode in which, as it were, the molecule is built up. First we have the "*Glyptic*" (γλυπτος, carved) formulæ of Dr. Hofmann, which is the simplest and perhaps the most perfect mode of representing molecular constitution. These consist of wooden balls, which may be made about two inches in diameter, and pierced with as many holes as the bonds of the atoms which they are intended to represent. The balls are usually coloured,—blue for hydrogen, red for oxygen, green for chlorine, white for nitrogen, and so on. They may be connected by pegs of wood, or short pieces of small brass gas-piping. These pegs represent the bonds or lines of force admirably. Bent pieces of piping, and short pieces of vulcanized tubing for connectors, will be found useful accessories.

Graphic formulæ are the pictorial representations of glyptic formulæ, in which atoms are shown by circles, and bonds by connecting lines. They are used when glyptic formulæ would be too cumbersome from the great complexity of some molecules, and may be made to gradually supersede them. Dr. Crum Brown introduced this form of formula as an aid to teaching, which has been adopted by Dr. Frankland mula SO_4H_2; this is an *empirical formula*, for it gives us no idea as to the mode whereby the atoms of sulphur, oxygen, and hydrogen are linked together; it merely gives us the sum-total of the atoms contained in the body, and enables us to calculate therefrom the respective weights of the elements present. *Constitutional* or *rational formulæ* are the symbolical representatives of the glyptic and graphic formulæ described above; they exhibit at the same time the number of atoms and weight of the elements, and also their mode of arrangement in the molecule. The following are the rules for formulating employed in this work:—

1. The first symbol in the formula is printed in thick type, unless the element be a monad.
2. The symbols of the elements or radicals whose available bonds are all directly united to the first symbol are made to follow it on the same line.
3. The element having the greatest number of bonds will, as a rule, be placed first.

Thus, sulphuric acid, the empirical formula of which is SO_4H_2, or H_2SO_4, when represented constitutionally becomes $\mathbf{S}O_2Ho_2$, which shows that the hexad sulphur atom is combined with the four bonds of the two oxygen atoms, and also with the two bonds of the two atoms of hydroxyl. This formula is therefore the exact representation of the graphic formula of the same substance which has already been given, p. 29.

A bracket is used for expressing chemical combination between two or more elements placed perpendicularly with regard to each other, and next to the bracket. Thus in—

$$\left\{\begin{array}{l}\mathbf{N}O_2\\ Cao''\\ \mathbf{N}O_2\end{array}\right.$$

the two nitrogen atoms have four of their bonds united to the four bonds of oxygen placed in the same line, and the remaining bond of the pentad nitrogen is fixed by a bond of the radical Cao″ placed perpendicularly with respect to it.

When the *b* hexads, p. 22, are found in combination as tetrads, the molecule almost invariably contains two atoms of the metal, which are linked together by one of their bonds. The tetrad aluminium is found in no other sort of combination, *e.g.*,—

$$\left\{\begin{array}{l}\mathrm{CrCl_3}\\\mathrm{CrCl_3},\end{array}\right. \quad \left\{\begin{array}{l}\mathrm{AlF_3}\\\mathrm{AlF_3},\end{array}\right. \quad \left\{\begin{array}{l}\mathrm{FeCl_3}\\\mathrm{FeCl_3},\end{array}\right. \quad \left\{\begin{array}{l}\mathrm{MnO}\\\mathrm{MnO}\end{array}\right.\!\mathrm{O}.$$

Three of the bonds of the metal atom are engaged in the horizontal line, and the remaining bond connects it with the other metal atom whose symbol is placed perpendicularly and next the bracket. The same kind of combination is exhibited in one line by using three atomicity marks to the right of the symbol and one to the left, thus:—

$$'\mathrm{Cr}'''_2\mathrm{Cl}_6,\ '\mathrm{Al}'''_2\mathrm{F}_6,\ '\mathrm{Fe}'''_2\mathrm{Cl}_6,\ '\mathrm{Mn}'''_2\mathrm{O}_3.$$

Metals of this class have been called pseudo-triads,

Constitutional Formulæ of some of the most Common Salts.

These formulæ are arranged so that the element represented by the left-hand symbol engages all the active atomicities of the other elements and compound radicals following upon the same line, and also all the spare atomicities of any element whose symbol is placed immediately above or below it. M represents the metal, compound radical, or basylous element. Mo, in the column of monads, represents a monad compound radical, such as Ko, Nao, &c. Mo″, in the diad column, represents a diad compound radical, such as Zno″ (ZnO_2), Feo″ (FeO_2). $(M_2O_6)^{vi}$ in the tetrad column represents a hexad radical; it may also be written in the abbreviated form M_2o^{vi}. In the formulæ of acids Ho represents the monad compound radical hydroxyl. Salts of pentads and hexads are rare.

Symbol of Acid.	Name of Acid.	Name of Salts.	Of Monads.	Of Diads.	Of Triads.	Of Tetrads.
	Monobasic Acids.					
$\left\{\begin{array}{l}OBr\\OHo\end{array}\right.$	Bromic	Bromates	$\left\{\begin{array}{l}OBr\\OMo\end{array}\right.$	$\left\{\begin{array}{l}OBr\\O\\Mo''\\O\\OBr\end{array}\right.$		
HBr	Hydrobromic	Bromides	MBr	MBr_2	MBr_3	MBr_4 $'M'''_2Br_6$
$\left\{\begin{array}{l}OCl\\OHo\end{array}\right.$	Chloric	Chlorates	$\left\{\begin{array}{l}OCl\\OMo\end{array}\right.$	$\left\{\begin{array}{l}OCl\\O\\Mo''\\O\\OCl\end{array}\right.$	.	

Symbol of Acid.	Name of Acid.	Name of Salts.	Of Monads.	Of Diads.	Of Triads.	Of Tetrads.
	Monobasic Acids (*continued*).					
HCl	Hydrochloric ……	Chlorides ………	MCl	$\mathbf{M}Cl_2$	$\mathbf{M}Cl_3$	$\mathbf{M}Cl_4$ $'\mathbf{M}'''_2Cl_6$
$\mathbf{O}ClHo$	Chlorous …………	Chlorites ………	$\mathbf{O}ClMo$	$\left\{\begin{matrix}\mathbf{O}Cl\\Mo''\\\mathbf{O}Cl\end{matrix}\right.$		
HF	Hydrofluoric………	Fluorides ………	MF	$\mathbf{M}F_2$	$\mathbf{M}F_3$	$\mathbf{M}F_4$ $'\mathbf{M}'''_2F_6$
$ClHo$	Hypochlorous ……	Hypochlorites …	$ClMo$	$MoCl_2$		
$\left\{\begin{matrix}\mathbf{O}I\\\mathbf{O}Ho\end{matrix}\right.$	Iodic ………………	Iodates …………	$\left\{\begin{matrix}\mathbf{O}I\\\mathbf{O}Mo\end{matrix}\right.$	$\left\{\begin{matrix}\mathbf{O}I\\O\\Mo''\\O\\\mathbf{O}I\end{matrix}\right.$		
HI	Hydriodic…………	Iodides …………	MI	$\mathbf{M}I_2$	$\mathbf{M}I_3$	$\mathbf{M}I_4$ $'\mathbf{M}'''_2F_6$
$\mathbf{N}O_2Ho$	Nitric………………	Nitrates…………	$\mathbf{N}O_2Mo$	$\left\{\begin{matrix}\mathbf{N}O_2\\Mo''\\\mathbf{N}O_2\end{matrix}\right.$	$\mathbf{N}_3O_6Mo''$	$\mathbf{N}_6O_{12}('M'''_2O_6)^{vi}$
$\mathbf{N}OHo$	Nitrous ……………	Nitrites…………	$\mathbf{N}OMo$	$\left\{\begin{matrix}\mathbf{N}O\\Mo''\\\mathbf{N}O\end{matrix}\right.$		

$\left\{\begin{matrix}OCl\\O\\OHo\end{matrix}\right.$	Perchloric	Perchlorates	$\left\{\begin{matrix}OCl\\O\\OMo\end{matrix}\right.$	$\left\{\begin{matrix}OCl\\O\\O\\Mo''\\O\\O\\OCl\end{matrix}\right.$	
	Dibasic Acids.			Of Diads.	Of Tetrads.
$COHo_2$	Carbonic	Carbonates	$COMo_2$	$COMo''$	
CrO_2Ho_2	Chromic	Chromates	CrO_2Mo_2	CrO_2Mo''	
$SSOHo_2$	Hyposulphurous ...	Hyposulphites ...	$SSOMo_2$	$SSOMo''$	
$\left\{\begin{matrix}COHo\\COHo\end{matrix}\right.$	Oxalic	Oxalates	$\left\{\begin{matrix}COMo\\COMo\end{matrix}\right.$	$\left\{\begin{matrix}CO\\CO\end{matrix}\right.Mo''$	
SO_2Ho_2	Sulphuric	Sulphates	SO_2Mo_2	SO_2Mo''	$\left.\begin{matrix}SO_2\\SO_2\\SO_2\end{matrix}\right\}-('M'''_2O_6)^{vi}$
SH_2	Hydrosulphuric ...	Sulphides	SM_2	SM	$\left\{\begin{matrix}'M'''S\\'M'''S\end{matrix}\right.S$
$SOHo_2$	Sulphurous	Sulphites	$SOMo_2$	$SOMo''$	
	Tribasic Acids.				
$AsOHo_3$	Arsenic	Arsenates	$AsOMo_3$	$\left\{\begin{matrix}AsOMo''\\Mo''\\AsOMo''\end{matrix}\right.$	

	Name of Salts.	Of Monads.	Of Diads.	Of Tetrads.
s				
.....	Arsenites	$AsMo_3$	$\left\{\begin{array}{l}AsMo''\\Mo''\\AsMo''\end{array}\right.$	
.....	Borates	BMo_3		
.....	Phosphates	$POMo_3$	$P_2O_2Mo''_3$	$P_4O('M'''_2O_6)^{VI}_3$
ic)				
.....	Metaphosphates .	PO_2Mo	P_2O_4Mo''	
c)				
.....	Pyrophosphates .	$P_2O_3Mo_4$	$P_2O_3Mo''_2$	
:)				
us .	Hypophosphites .	POH_2Mo	$\left\{\begin{array}{l}POH_2\\Mo''\\POH_2\end{array}\right.$	
.....	Phosphites	$POHMo_2$	$POHMo''$	
ls.				
.....	Silicates...........	$SiMo_4$	$SiMo''_2$	$Si_3O_3('M''_2O_6)^{VI}$

	Of Monads.	Of Diads.	Of Triads.	Of Tetrads.
....	M_2O	MO	M_2O_3	MO_2
s............	MHo	MHo_2		$\left\{\begin{array}{l}MHo_4\\'M'''_2Ho_6\end{array}\right.$

onobasic. † Dibasic. ‡ Tetrabasic.

Equations.—Chemical changes are explained by equations, in which the elements or compounds concerned in the change are placed on the left of an = sign, with a + sign between the molecules, and the products of the reaction are placed to the right of the = sign, also with the sign + placed between their molecules. Equations in which single atoms of the elements or radicals occur are said to be *atomic equations;* they are not correct representations of chemical change, as by them single atoms are made to have a separate existence. They are correct only in cases where a single atom is also a molecule; *e. g.*, zinc, cadmium, mercury.

Molecular equations, which represent more correctly what takes place, are much to be preferred; the following are examples of the two kinds:—

Atomic Equations.

$$\mathrm{Na} + \mathrm{OH_2} = \mathrm{NaHo} + \mathrm{H}.$$
$$\mathrm{K} + \mathrm{HCl} = \mathrm{KCl} + \mathrm{H}.$$

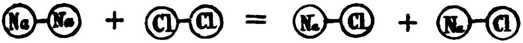
Molecular Equations.

$$\mathrm{Na_2} + 2\mathrm{OH_2} = 2\mathrm{NaHo} + \mathrm{H_2}.$$
$$\mathrm{K_2} + 2\mathrm{HCl} = 2\mathrm{KCl} + \mathrm{H_2}.$$
$$\mathrm{Zn} + \mathrm{OH_2} = \mathrm{ZnO} + \mathrm{H_2}.$$

Atomic and Molecular Combination.—In all examples of chemical combination hitherto given it has been essentially atomic, atom being linked to atom, atom uniting with atom. In all probability, where direct combination occurs, there is a breaking up or opening out of previously existing molecules: thus, when common salt, NaCl, is formed by burning sodium in chlorine, a molecule of sodium is acted on by a molecule of chlorine, and it is rather by an interchange of atoms than by direct combination that the product is formed; thus,—

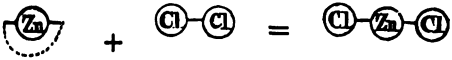
$$\mathrm{Na_2} + \mathrm{Cl_2} = 2\mathrm{NaCl}.$$

And it may be represented graphically thus,—



Molecule of sodium.	Molecule of chlorine.	Two molecules of sodic chloride.

And when the monatomic molecule of zinc is acted on by a molecule of chlorine, the latent bonds of the zinc are opened out and become active; thus,—



Molecule of zinc.	Molecule of chlorine.	Molecule of zincic chloride.

There are other cases also, though somewhat rare ones, in which two molecules, by the opening out of their latent bonds, unite to form one molecule; thus,—

$$N'''H_3 \quad + \quad HCl \quad = \quad N^{v}H_4Cl.^{*}$$

One molecule of ammonia.	One molecule of hydrochloric acid.	One molecule of ammonic chloride.

$$C''O \quad + \quad Cl_2 \quad = \quad C^{iv}OCl_2.$$

One molecule carbonic oxide.	One molecule chlorine.	One molecule phosgene gas.

There are, however, undoubted cases in which molecules unite without altering the active atomicity

APPARATUS AND MANIPULATION.

Before beginning the practical part of the science it may not be amiss if a few words of advice and caution be introduced here as to Apparatus and Manipulation.

Corks are of constant use, and they should be of the best kind. Before fitting them into apparatus they should be softened by rolling them under the foot in a piece of paper, so as to keep them free from dust.

When glass tubes have to be passed through the cork, holes should be made with a cork-borer rather smaller than the tube to be inserted, and nicely eased to the proper size with a rat-tail file. If the cork should prove porous, the exposed parts must be covered with a solution of sealing-wax in spirits of wine.

Glass tubing may be cut into the required lengths by making a deep scratch on the tube with a triangular file; the tube is then easily broken. The depth of the mark made by the file must depend on the thickness of the glass. If the tube has to be bent, this should be done before cutting it. To do this the tube must be heated gradually in the flame of a gas or spirit lamp, by turning it round in the flame, and moving it a little to the right and left. As soon as the tube becomes red-hot it can be easily bent into the required shape. A little practice with a few pieces of glass tube over a lamp will soon give the experimenter more knowledge than any written instructions.

Vulcanized india-rubber tubing, which can be purchased of any size, is used for connecting glass tubes. In connecting glass tubes by means of india-rubber, the ends of the tube should not be brought so close together as to interfere with the flexibility of the arrangement. If the india-rubber tubing does not fit the glass tubes perfectly airtight, it must be tied with *a piece of twine.*

Florence flasks will be found extremely useful, as they can be fitted up to suit a variety of purposes, and are very cheap. There are two sizes, holding one quart and half a pint respectively. They may be freed from the oil by boiling a little washing soda or ammonia in them.

Gas Jars.—Plain cylindrical jars should be from twelve to sixteen inches in length, and about two and a half to three inches in diameter, and about one-eighth of an inch in thickness. The open end should be ground, so as to be closed airtight with a piece of flat glass, which also should be ground. A little grease should be placed on the glass plate or edge of the gas jar, so as to insure perfect tightness. Stoppered jars usually have glass stoppers like bottles, and should be ground at the top, so as to be closed with a glass plate in the manner described for cylindrical gas jars. These jars are removed from the pneumatic trough by passing a plate or saucer under the bottom. The plate or saucer should always contain a little water, so as to prevent the access of air and the escape of the gas.

When substances have to be burned in a gas jar, they are introduced by means of a deflagrating spoon. This spoon should be fitted with a plate, so as to cover the aperture of the jar, and allow it to pass the required depth into the gas. In burning phosphorus, a piece not larger than a pea should be used, unless a large jar or globe be employed, when the piece may be larger, and it will give a much more brilliant light; it should be carefully dried by gently wrapping it in blotting-paper. The phosphorus may then be ignited by touching it with a red-hot wire and quickly introducing it to about two-thirds the depth of the jar.

In collecting gases over water, where heat is required, the heat should in all cases be applied gradually, and the first portions of gas allowed to escape, till all the common air in the apparatus has been expelled. At the close of the experiment the delivery pipe should be removed from the water to prevent its receding.

The pneumatic trough should be about two feet long, eighteen inches wide, and a foot deep, provided with a shelf about three inches below the edge of the trough. The water should always stand about one inch over this shelf. Holes are made in this shelf, over which the jars intended to receive the gas are placed. The delivery tube is brought under the open end of the jar, and as the gas rises in the jar the water is displaced. A spout should be provided to the trough, so as to allow the surplus water to pass into a pail. An ordinary gas jar is filled with air, and, as no two bodies can occupy the same space at the same time, it is necessary to expel this air before the jar can be filled with gas. This is done by immersing the jar sideways in the water of the pneumatic trough, turning it over with the open end downwards, and gently raising it to the shelf of the trough. If the stopper of the jar be airtight, the water, although above the level of that in the trough, will stand for any length of time in the jar. A shallow pail, or even a wash basin or bowl, forms a good substitute for the pneumatic trough.

Woulff's Bottles.—It is sometimes necessary to dry gases. This is effected by passing them through a series of Woulfi's bottles containing sulphuric acid, or passing the gas through a tube, from a foot to eighteen inches long and three quarters of an inch in diameter, filled with calcic chloride. Tubes in the form of a U are frequently used. The sulphuric acid and calcic chloride are called *desiccating agents*, because they deprive gases of water.

In the generation of carbonic oxide from *oxalic acid and sulphuric* acid, carbonic anhydride and carbonic oxide are evolved simultaneously. The presence of the former interferes with a careful examination of the carbonic oxide. This substance must be removed; and this is effected by passing the gases through a strong solution of potassic hydrate (caustic potash) or lime water, which retains the anhydride. By this method pure carbonic oxide is obtained, which is passed into gas jars filled with water at the pneumatic trough.

When no heat is required to liberate gases or gaseous compounds, a two-necked Woulff's bottle may be used. It should be fitted up with two glass tubes. The delivery tube, from which the gas escapes, may be about a quarter of an inch in diameter, and pass just through the cork of the bottle; another tube, or better still, an acid funnel of suitable length passing to within a quarter of an inch of the bottom of the bottle, is used for pouring in acid or water. This saves the trouble of taking out the cork, which must be done every time fresh acid or water is added, unless the bottle be provided with two tubes in the manner described. The materials the same as gases of nearly the same sp. gr. of air. If the gases are lighter than common air, as hydrogen and ammonia, or heavier, as chlorine and hydrochloric acid, the pneumatic trough may be dispensed with. But gases of nearly the same specific gravity as common air must be collected over water. The same generating apparatus as that just described will do, provided the flask or bottle will stand heat if required. The delivery tube must pass to the bottom of the jar it is intended to fill with gas. This is often done by blocking the jar up with pieces of wood of different thicknesses. The top of the jar may be covered with a piece of cardboard, with a hole for the delivery tube to pass through. As the gas collects at the bottom it gradually expels all the air, exactly the same as if water were poured into the gas jar. (See *Fig.* 1.) In raising bottles or other apparatus let the blocking be firm, the largest and heaviest pieces at bottom. When the gases are lighter than common air, the receiver may be supported by a retort-stand, the delivery tube passing upwards to the extremity of the vessel intended *to receive the gas.* When large quantities of

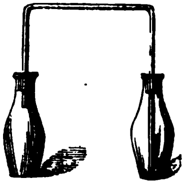


FIG. 1.

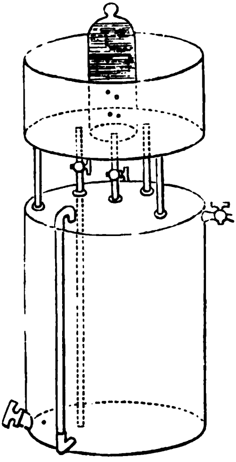


FIG. 2.

gas are required a gasometer is used. The most convenient is the one known as Pepys' gasholder. (See *Fig.* 2.) Five minutes' examination of one of these gasholders will give the student a much better idea of its action than any written description. The most careful explanation of any piece of mechanism or apparatus must fall short of actual observation.

In chemistry the terms *solution*, *soluble*, and *solvent* are frequently employed. Many solid bodies, when placed in water, dissolve. Salt and sugar dissolve in water. These substances are said to be soluble in water; and the water is called the solvent or menstruum. The more completely a solid is reduced to powder the easier it dissolves. If sugar be placed in alcohol it does not dissolve. Here, then, we have an example, which may be multiplied to a great extent, of a body soluble in one liquid and insoluble in another. A solution may be either mechanical or chemical. The sugar in water is a case of mere mechanical solution; no chemical change has taken place, and, by evaporation, the sugar can be obtained in its original condition. Salt dissolved in water is another illustration of the same kind. In chemical solutions the solid and liquid react on each other, forming a new compound, from which the original solid and fluid cannot be obtained by evaporation or any mechanical operation. When chalk is dissolved in hydrochloric acid, or copper dissolved with nitric acid, we have examples of chemical solution. A liquid can only dissolve a certain quantity of a solid substance. This quantity will vary according to the solid and liquid employed. When the solvent or liquid contains as much of the solid matter as it is capable of dissolving, it is said to be *saturated*, as all additional quantities of the solid fall to the bottom.

The solubility of most substances is increased by an increase of temperature.

100 pts. of	water at	15·5°C.	dissolve	11	pts. of potassic sulphate.
,,	,,	100°	,,	25	,, ,,
,,	,,	15·5°	,,	32	pts. of dry magnesic sulphate.
,,	,,	100°	,,	74	,, ,,

Common salt is about as soluble in hot water as in cold. Some bodies attain their maximum solubility at a particular temperature. The solubility of dry sodic sulphate is at its maximum at a temperature of 34°; but in sodic nitrate we have the order reversed,—water, at 0°, dissolves 80 parts; at 10° only 23 parts; but from this point the solubility increases up to 200 parts at a temperature of 100° C.

Substances dissolved in alcohol, ether, oils, and water are called simple solutions. The most important solvent is water; and when the substance is soluble in this liquid no other is used.

Elevation of temperature accelerates all chemical solutions; heat promotes the action of bodies on each other. When bodies enter into chemical combination they unite in definite and fixed proportions, irrespective of temperature. Solutions are made either in test-tubes, Florence flasks, evaporating dishes, or beakers, the size being regulated by the quantity of the substance to be dissolved.

Filters and Filtration.—Filtering paper should be of uniform texture, and sufficiently porous to admit a free passage of the fluid. Filtration is purely a mechanical operation. The object is to separate the solid from the liquid; the former collecting on the filter, and the latter passing through, which is called the *filtrate.* The filtering paper is supported by a glass funnel, and the funnel may be supported by one of the rings of the retort-stand. (See *Fig.* 3.) A filter is easily made by folding the paper in such a manner as to bring the four corners together, and by opening one corner from the other three, a conical cavity is formed. The filter may now be placed in the funnel, taking care that it is the proper size. It should fit nicely against the sides of the funnel, and the substance to be filtered added by gently pouring it on the side. If projected violently into the centre of the filter, the paper is liable to be broken. Nothing short of actually doing the thing will give the proper knowledge. It would be just as reasonable to expect a man to learn the trade of a joiner by reading books on the construction of tools, as to expect a person to learn chemistry without proper apparatus; but this apparatus need not be expensive, if the student have (as he ought to have) some inventive and constructive power.

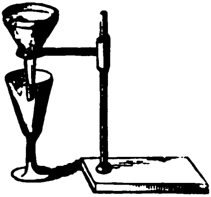


FIG. 3.

The following list of apparatus suitable for perform-

Brought forward	11	2
1 Bunsen burner, single jet	1	6
Or—		
1 spirit lamp, 4 oz., 1s., where there is no gas.		
½ dozen London fireclay crucibles, round, 3¼ in.	0	6
½ dozen covers for crucibles	0	9
1 Berlin porcelain crucible and cover, ½ oz.	0	8
1 pair of brass crucible tongs, 6 in.	1	0
1 deflagrating jar, 9 × 4½ in.	2	0
2 gas cylinders, 8 × 1½ in.	2	0
1 gas cylinder, 12 × 2 in.	1	6
1 deflagrating spoon	1	0
1 beehive shelf, 4 in.	0	8
Vulcanized india-rubber tubing, ¼ in. bore, 3d. per foot	0	6
Ditto, ⅜ in. bore	0	6
1 Wedgwood-ware mortar and pestle, 4 oz.	1	2
Retorts, two 4-oz., one 8-oz., stoppered	2	1
2 lbs. glass tubing, mixed bore	2	8
½ lb. glass rod for stirrers, &c.	0	8
½ dozen test-tubes, 6 × ½ in.	0	6
½ dozen " 6 × ¾ in.	0	7
1 dozen 4 oz. bottles, narrow-mouthed, stoppered	3	6
3 12 oz. " " "	1	9
1 dozen 6 oz. wide-mouthed bottles for corks	2	0
1 Woulff's bottle, two necks, pint size	1	6
1 test-tube stand to hold six	0	9
Piece of wire gauze to put under flasks, &c., when being heated over gas lamp	0	3
Piece each of platinum wire and foil	0	9
1 Black's blowpipe	0	9
1 retort-stand, 16 in. high, 3 rings	3	6
An assortment of good corks.		
Total . . £2	6	2

The above set of apparatus, or any part of it, may be had of George *Jarmain, 80,* Northgate Street, Huddersfield.

The following apparatus will be found extremely useful, but they are not indispensable:—

	s.	d.
1 Mohr's burette, 50 c.c. in $\frac{1}{8}$ c.c.	5	0
1 flask to hold 1 litre	3	0
1 pipette to deliver 100 c.c.	1	9
1 gas receiver with brass cap, &c.	5	6
1 brass stopcock, male screws	3	0
1 Bunsen's eudiometer, 500 m.m. in 500 divisions	6	0
1 centigrade thermometer, with glass scale, 200°	4	0
1 universal wooden support	5	0
1 Mackintosh bag, 18 × 13 in., fitted with brass screw	10	0
1 Herapath's blowpipe, to fit Bunsen's burner	2	6
1 box of scales and weights, oak box	2	6
1 iron tripod stand	1	0
½ dozen test-glasses, cylindrical, 6 oz.	3	6
1 wash-bottle with blowing tube	1	0
1 white earthenware tile	0	6
1 sheet-iron sand bath, 4 in.	0	4
2 Berlin porcelain evaporating basins, 8 oz. and 16 oz.	2	7
Total	£2 17	2

The tyro in chemistry is often at a loss to know what materials he should purchase, and the quantity of each kind; a list of substances is therefore appended, from which he may himself prepare most of the compounds named in this treatise. A more extensive list will also be found in the Appendix.

LIST OF USEFUL CHEMICALS.

Systematic Name.	Formula.	Shop or Commercial Name.	Price.	Suitable quantity to purchase.	Tradesman who supplies the article.
Sulphuric acid	SO_2Ho_2	Oil of vitriol.	2d. per lb.	1 lb.	Chemist and druggist.
Hydrochloric acid	HCl	Muriatic acid.	2d. per lb. 8d. per lb. pure	1 lb.	Do.
Nitric acid	NO_2Ho	Aquafortis.	4d. per lb. 1s. per lb. pure	1 lb.	Do.
*Oxalic acid	$\left\{\begin{matrix}COHo\\COHo\end{matrix}\right.$		1s. 2d. per lb.	¼ lb.	Do. or drysalter.
Zinc	Zn	Zinc clippings.	1½d. per lb.	1 lb.	Tinner.
Sodium	Na		1s. per oz.	½ oz.	Chemist and druggist.
Iron	Fe	Iron filings and borings.			Any machine shop.
Copper	Cu	Copper turnings.	1s. 4d. per lb.	¼ lb.	Do.
Manganic oxide	MnO_2	Black oxide of manganese.	3d. per lb.	1 lb.	Chemist and druggist.
Mercuric oxide	HgO	Red precipitate.	3d. per lb.	1 oz.	Do.
*Potassic chlorate	$\left\{\begin{matrix}OCl\\OKo\end{matrix}\right.$	Chlorate of potash.	1s. 6d. per lb.	½ lb.	Do.
,, bichromate	$\left\{\begin{matrix}CrO_2Ko\\O\\CrO_2Ko\end{matrix}\right.$	Bichromate of potash.	6d. per lb.	½ lb.	Drysalter or do.

[redacted] *of Useful Chemicals (continued).*

[redacted]	Shop or Commercial Name.	Price.	Suitable quantity to purchase.	Tradesman who supplies the article.
[redacted]	Subcarbonate of potash.	8d. per lb.	¼ lb.	Chemist and druggist.
[redacted]	Iodine of potassium.	1s. 2d. per oz.	½ oz.	Do.
[redacted]	Saltpetre.	7d. per lb.	½ lb.	Do.
[redacted]	Yellow prussiate of potash.	1s. 3d. per lb.	¼ lb.	Drysalter or druggist.
[redacted]	Bromide of potassium.	10d. per oz.	½ oz.	Chemist and druggist.
[redacted]	Permanganate of potash.	1s. per oz.	1 oz.	Do.
[redacted]		4d. per oz.	1 oz.	Do.
[redacted]	Chloride of barium.	2d. per oz. pure.	1 oz.	Do.
[redacted]		2d. per lb.	1 lb.	Drysalter or druggist.
[redacted]		4d. per lb.	¼ lb.	Chemist and druggist.
[redacted]	Slaked lime.			
[redacted]		1d. per oz.	1 oz.	Do.
[redacted]	Carbonate of soda.	Howard's 8d., others 4d. pr. lb	¼ lb.	Do.
[redacted]	Washing soda.	1d. per lb.	1 lb.	Grocer, &c.

List of Useful Chemicals (continued).

Systematic Name.	Formula.	Shop or Commercial Name.	Price.	Suitable quantity to purchase.	Tradesman who supplies the article.
Calcic fluoride . .	CaF_2	Fluor spar.	4d. per lb.	1 lb.	Chemist and druggist.
,, carbonate . .	$COCao''$	Marble, chalk, or calc spar.			Marble mason.
Ferrous sulphate . .	SO_2Feo''	Copperas.	2d. per lb.	1 lb.	Drysalter or druggist.
Argentic nitrate . .	NO_2Ago	Nitrate of silver, crystallized.	4s. per oz.	¼ oz.	Chemist and druggist.
Ferrous sulphide . .	FeS	Sulphide of iron.	6d. per lb.	1 lb.	Do.
Carbon	C	Charcoal. Bone-black.			
Sulphur . . .	S	Roll sulphur.	3d. per lb.	1 lb.	Do.
*Iodine . . .	I		1s. 4d. per oz.	¼ oz.	Do.
*Tin	Sn	Grain tin.	1s. 6d. per lb.	¼ lb.	Drysalter or tinner.
Hydric-disodic phosphate . . .	$POHoNao_2$	Phosphate of soda.	2d. per oz.	1 oz.	Chemist and druggist.
Antimony . . .	Sb		2d. per oz.	1 oz.	Do.
Antimonious sulphide .	Sb_2S_3	Black antimony.	8d. per lb.	¼ lb.	Do.
*Bismuthous nitrate .	N_3O_6Bio'''	Trisnitrate of bismuth.	1s. 4d. per oz.	¼ oz.	Do.
Strontic nitrate . .	$\left\{\begin{matrix}NO_2\\Sro''\\NO_2\end{matrix}\right.$	Nitrate of strontia.	2d. per oz.	1 oz.	Do.

ful Chemicals (*continued*).

op or Commercial Name.	Price.	Suitable quantity to purchase.	Tradesman who supplies the article.
Magnesium ribbon.	4d. per yard.	1 yard.	Chemist and druggist.
psom salts.	4d. per lb.	¼ lb.	Do.
uicksilver.	2s. 6d. per lb.	½ lb.	Do.
lue vitriol.	5d. per lb.	1 lb.	Drysalter or do.
ichloride of platinum.	4s. per dram.	½ dr.	Chemist and druggist.
craps of old ad, melted.			
gar of lead.	8d. per lb.	¼ lb.	Do.
Red-lead.	4d. per lb.	¼ lb.	Do. or oilman.
Nitrate of cobalt.	2s. per oz.	½ oz.	Chemist and druggist.
ulphate of nickel.	1s. per oz.	½ oz.	Do.
	10d. per pint.	1 pint.	Do.

ked * fluctuate in price considerably.

NON-METALLIC ELEMENTS.

MONADS.

Hydrogen—H_2.

Atomic wt. and sp. gr. = 1. Molecular weight = 2. Molecular volume, normal. 1 litre weighs 1 crith.* Atomicity, I. The standard of comparison.

This gas exists largely in a state of combination in the organic and inorganic kingdoms; it is a constituent of water and of most substances used either for fuel or illumination; it is an essential constituent of all acids—hydrochloric, HCl; hydrobromic, HBr; hydriodic, HI; hydrosulphuric, SH_2; sulphuric, SO_2Ho_2, &c. Most animal and vegetable substances contain it; it is also found in many minerals. It is only met with in a free state in nature in volcanic gases.

Preparation.—It is prepared in a variety of ways:—

(1.) By passing a small piece of sodium or potassium into an inverted tube filled with water, and standing on the pneumatic trough:—

$$Na_2 + 2OH_2 = \underset{\text{Sodic hydrate.}}{2ONaH} + H_2.$$

The sodium displaces half the hydrogen from the portion of water acted on, whereby it is converted into sodic hydrate, which dissolves as rapidly as it is formed. The displaced hydrogen collects in the upper part of the tube. The sodium should be folded in a piece of blotting-paper and rapidly introduced into the tube. This experiment is attended with some danger unless skilfully performed.

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* See paragraph on the crith and its uses.

(2.) By passing steam through a red-hot iron tube filled with nails or iron borings. The iron displaces the hydrogen, which escapes as gas, thus:—

$$3Fe + 4OH_2 = \underset{\text{Magnetic oxide.}}{Fe_3O_4} + 4H_2.$$

(3.) By electric decomposition of water, as will be described when treating of oxygen:—

$$2OH_2 = O_2 + 2H_2.$$

The water breaks up into its elements, oxygen and hydrogen, which may be collected separately.

(4.) By adding diluted sulphuric acid to zinc clippings, in a bottle provided with a funnel tube passing to its bottom, and a tube for conducting the evolved gas to the pneumatic trough:—

$$Zn + \underset{\text{Sulphuric acid.}}{SO_2Ho_2} = \underset{\text{Zincic sulphate.}}{SO_2Zno''} + H_2.$$

The zincic sulphate remains behind in the bottle, and if the liquid be filtered and set aside in a shallow dish the salt will crystallize in a few hours. The acid

(6.) By passing steam through a small platinum pipe, which is heated by the oxyhydrogen blowpipe flame. The intense heat overcomes the affinity of the oxygen for the hydrogen.

Properties.—Hydrogen at all known temperatures and pressures is a transparent and colourless gas, and when pure is tasteless and without odour. It is the lightest substance known. A litre of hydrogen at the normal temperature and pressure only weighs ·0896 gramme, whilst a litre of air weighs nearly 14¼ times as much, and one of oxygen 16 times as much under like conditions.

Hydrogen, from its extreme lightness, was formerly used for filling balloons; its use is now superseded by employing coal gas and larger balloons.

The property of diffusion, which is common—in a greater or less degree—to all gases and vapours, is well seen in the following experiment with hydrogen and carbonic anhydride, two bodies of very different specific gravities. If two bottles, one containing hydrogen and the other carbonic anhydride, be joined together with a glass tube about a foot long and 3-8ths of an inch in diameter, the hydrogen in the upper bottle and the carbonic anhydride in the lower, we shall find, after a few hours, that the carbonic anhydride (contrary to its gravity) and the hydrogen have uniformly mixed throughout. This diffusive power of gases is of great importance in the economy of the universe. Were it not for it we should have the different gases and vapours of the atmosphere arranging themselves on the earth's surface according to their respective weights. This diffusive power of hydrogen renders it impossible to keep it for any length of time in bladders, or the best constructed gasholders, and when mixed with air from any leakage it explodes when ignited.

Hydrogen is not a supporter of combustion. If a lighted taper be plunged into a jar of hydrogen it is extinguished; the hydrogen burns only at the top of the jar, where it comes in contact with the oxygen of the air. A jet of chlorine, oxygen, or air, will, however, burn in an atmosphere of hydrogen. Hydrogen burns with a pale blue flame, without much light, but with intense heat. The temperature may be increased by mixing hydrogen and oxygen, in the proportion of two volumes of the former to one of the latter, and burning it from a safety jet. This is called the oxyhydrogen blowpipe. It is best to use two separate gasholders for this experiment. When hydrogen is mixed with common air or pure oxygen it is highly explosive. Great care should therefore be taken in preparing it to allow the first portions of gas to escape before applying a light. Soap-bubbles and india-rubber balloons, blown with hydrogen by means of a common pipe attached to a bladder provided with a stopcock to regulate the emission of the gas, rise in the air with considerable velocity. If a lighted taper be applied to the bubble as it ascends, a slight explosion piece of spongy platinum be brought near a jet of hydrogen, it is at once kindled, and the platinum becomes red-hot. The hydrogen and atmospheric oxygen are so condensed in this experiment that they are brought within the sphere of attraction necessary for combination. The same takes place with a mixture of the two gases.

The chemical relations of hydrogen assign to it the character of a metal, and many efforts have been made to establish its metallic character, but all attempts to liquefy or solidify it by the usual modes have hitherto failed. Professor Graham, of Her Majesty's Mint, however, has recently obtained an alloy of hydrogen and palladium, having a composition closely approaching one atom each of the two elements.* "Hydrogenium," as Prof. Graham calls the combined hydrogen, "acts in this case like a metal, and thus the opinion held by some scientific men, that hydrogen constitutes the vapour of a metal, receives confirmation. The sp. gr. of hydrogenium as contained in the alloy was found to be 1·95." The alloy was obtained by using a palladium pole for a galvanic water decomposer, when the hydrogen was absorbed by the palladium instead of being given off. Palladium was found to absorb 900 times its bulk of hydrogen.

Nascent Hydrogen.—When an elementary body is in the act of being liberated from a state of combination it is said to be in a nascent condition, and in this state the activity of its properties is greatly enhanced, and it is in consequence more disposed to enter into fresh combinations than when it has become entirely free. To account for this phenomenon it may be assumed that the elementary atoms are liberated singly, but that they immediately combine with each other to form elementary molecules, in case nothing else be present with which they can combine. Nascent hydrogen shows this enhanced power of combination in an eminent degree; by virtue of this power it acts as a *reducing agent*,—that is, it will combine with and thus remove the whole or a portion of the oxygen or chlorine, &c., which bodies contain, and thus it *reduces* them to a lower state of combination. Many cases of this kind will be noticed in this work, for nascent hydrogen is a favourite reducing agent. The following will serve at present as an illustration:—

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* *Chemical News*, August 20, 1869.

If hydrogen be passed through a solution of ferric chloride, Fe_2Cl_6, for any length of time, it will not alter its composition, for free hydrogen cannot reduce Fe_2Cl_6 in solution; if, however, the hydrogen be generated in the solution itself, the nascent hydrogen combines with two chlorines, and thus reduces ferric to ferrous chloride, thus:—

$$\underset{\text{Ferric chloride.}}{Fe'''_2Cl_6} + 2HCl + Zn = \underset{\text{Zincic chloride.}}{ZnCl_2} + \underset{\text{Ferrous chloride.}}{2Fe''Cl_2} + 2HCl.$$

chloride, NaCl). It is never found in nature in its free or uncombined state. It is often found among the gases evolved from volcanoes as hydrochloric acid, HCl.

Preparation.—It is prepared by mixing, to the consistency of cream, strong hydrochloric acid and manganic oxide (black oxide of manganese). Care should be taken that every part is well mixed, so that no dry place appears inside the retort or flask, or fracture is most likely to result. A gentle heat should now be applied, when the gas comes over in abundance. It should be collected over warm water, as cold water absorbs about twice its volume; or it may be collected by displacement (see p. 43). The reaction is easily understood:—

$$(1.)\ MnO_2 + 4HCl = \underset{\text{Manganic chloride.}}{MnCl_4} + 2OH_2.$$

This manganic chloride under the influence of heat decomposes, thus:—

$$(2.)\ Mn^{\text{iv}}Cl_4 = \underset{\substack{\text{Manganous}\\\text{chloride.}}}{Mn^{\prime\prime}Cl_2} + Cl_2.$$

Instead of the above, a mixture of common salt, NaCl, sulphuric acid, and manganic oxide may be used; the following reaction then occurs:—

$$MnO_2 + \underset{\substack{\text{Sodic}\\\text{chloride.}}}{2NaCl} + 2SO_2Ho_2 = \underset{\substack{\text{Manganous}\\\text{sulphate.}}}{SO_2Mno^{\prime\prime}} + \underset{\substack{\text{Sodic}\\\text{sulphate.}}}{SO_2Nao_2} + 2OH_2 + Cl_2.$$

This gas produces great irritation in the air passages even when breathed in small quantities, and should therefore be prepared in the open air, or in a well-ventilated room.

Properties.—Chlorine is a yellowish green gas at ordinary temperatures, but under a pressure of 4 atmospheres at 15·5° C. it is reduced to a limpid liquid, having a sp. gr. of 1·33. Chlorine can be obtained in combination with every elementary body.

Chlorine has little affinity for oxygen; it combines readily, however, with hydrogen and the metals. If a lighted taper be immersed in a jar of the gas it burns with a dull red light, with an abundant evolution of carbon, in the form of smoke. If a piece of blotting-paper be soaked in turpentine, and placed in a jar of chlorine, it bursts into flame. In both these cases the chlorine combines with the hydrogen to form hydrochloric acid, to the exclusion of the carbon, which falls as soot—a very good illustration of the superior affinity of chlorine for hydrogen.

Phosphorus combines with chlorine spontaneously, but the flame is very feeble. Copper leaf, powdered antimony, and arsenic, all undergo spontaneous combustion in chlorine. If equal volumes of hydrogen and chlorine be introduced into a suitable apparatus, and an electric spark sent through the mixture, combination takes place with great violence. The direct rays of the sun or the approach of a lighted taper to tion, if it be poured into test-glasses containing infusions of litmus, purple cabbage, turmeric, carmine, cochineal, &c., the colours of these substances will be removed. When bleaching is practised on a large scale a combination with lime is used, called *bleaching powder*. The outline of this process can be easily imitated. The bleaching powder is mixed with warm water and the solution filtered. Any piece of coloured calico may be taken, providing its tint is due to some organic colouring matter. It may then be washed in the filtered solution. The liquid bleaches but slightly; but if it be removed from the bleaching bath, and rinsed in a bath of dilute sulphuric acid, and this process carried on a few times alternately, the whole of the colour is removed. The chlorine, when in combination with oxygen and calcium, is comparatively inactive; but when dipped in dilute acid the combination is broken up, and the chlorine in the presence of water is set free, and manifests its bleaching powers. An experiment illustrating this property may be performed in the following manner:—A piece of cotton, dyed with madder, or any organic colouring matter, may have any pattern or device stamped upon it, or painted with a mixture of gum and tartaric acid. The cloth is dried, and afterwards rinsed in a hot solution of bleaching powder. A white pattern on a coloured ground is developed, in consequence of the acid mixed with the gum liberating the chlorine, and destroying the colour in those parts upon which the pattern was printed. The other parts of the cloth are not sensibly affected.

The property of bleaching is due to the great affinity of chlorine for hydrogen; it decomposes vegetable and animal substances, forming with them colourless compounds containing chlorine; in many cases the bleaching effect is no doubt due to the liberation of oxygen from water, the presence of which is essential:—

$$2OH_2 + 2Cl_2 = 4HCl + O_2.$$

The *nascent* oxygen acts on the colouring matter, and forms a colourless compound with it.

Chlorine no doubt owes its potent disinfecting and deodorizing effects to the same properties.

Hydrochloric Acid—HCl.

Molecular wt.=36·5. Molecular volume, normal. 1 litre weighs 18·25 criths.

Preparation.—As already seen, this acid may be prepared by exploding a mixture of equal volumes of chlorine and hydrogen, no alteration of volume taking place:—

$$H_2 + Cl_2 = 2HCl.$$

Hydrochloric acid gas is commonly prepared by heating in a properly arranged flask a mixture of common salt, a little water, and sulphuric acid. It must be collected by displacement, or over mercury. A

$$\begin{array}{cccc} NaCl + & SO_2Ho_2 = & HCl + & SO_2HoNao. \\ \text{Sodic} & \text{Sulphuric} & \text{Hydrochloric} & \text{Hydric-sodic} \\ \text{chloride.} & \text{acid.} & \text{acid.} & \text{sulphate.} \end{array}$$

Hydrochloric acid is made on the large scale in iron retorts; the product is impure, containing ferric chloride, sodic chloride, arsenious chloride, sulphuric and sulphurous acids, and chlorine. This crude acid is called muriatic acid, or spirit of salt, and is largely used as a solvent for tin for the use of the dyer and calico-printer. Hydrochloric acid is also a large by-product in the manufacture of soda ash.

Properties.—It is a colourless gas, fuming when in contact with the atmosphere, and is very soluble in water. At 15° C. water absorbs 458 times its bulk, and has then a sp. gr. of 1·21. The gas contains equal volumes of hydrogen and chlorine, which may be proved as follows:—

(*a*) *By Synthesis:* 1. One volume of hydrogen and one volume of chlorine exploded yield 2 volumes of hydrochloric acid.

(*b*) *By Analysis:* 2. A volume of hydrochloric acid gas in which metallic potassium is made to burn is reduced to half its bulk, and the residue is hydrogen.

3. Solution of HCl when decomposed by the galvanic current, after the lapse of some time, yields equal volumes of hydrogen and chlorine.

By the action of metals, oxides, hydrates, &c., on hydrochloric acid, chlorides are obtained thus:—

$$\begin{aligned} Zn + 2HCl &= ZnCl_2 + H_2. \\ CaO + 2HCl &= CaCl_2 + OH_2. \\ ONaH + HCl &= NaCl + OH_2. \\ COCao'' + 2HCl &= CaCl_2 + OH_2 + CO_2. \end{aligned}$$

A mixture of hydrochloric and nitric acid is called *aqua regia.* This substance is a powerful solvent.

Test for HCl.—A solution of $\mathbf{N}O_2Ago$, added to a solution containing HCl, or a chloride, will give a white curdy precipitate, which is insoluble in $\mathbf{N}O_2Ho$, but soluble in AmHo; *e. g.*—

$$NaCl + \mathbf{N}O_2Ago = \underset{\text{Insoluble precipitate.}}{AgCl} + \mathbf{N}O_2Nao.$$

DIAD ELEMENTS.

Oxygen—O_2.

Atomic weight and sp. gr. = 16. Molecular weight, 32, Molecular volume, normal. 1 litre weighs 16 criths. Atomicity, II.

Oxygen gas was discovered by Priestley in 1774, who called it *dephlogisticated air.* It was also discovered about the same time by Scheele, a Swedish chemist, who called it *empyreal air* and *vital air.* The term oxygen was given to this gas by Lavoisier some time after its discovery.

Occurrence in Nature.—This element is largely

1. By exposing the red oxide of mercury to heat in small glass retort or hard glass test-tube. In this bstance the affinity which holds the oxygen to the ercury is so feeble, that exposure to heat is sufficient cause a separation into oxygen gas and metallic ercury. The mercury collects in small globules in e cold part of the retort, and the oxygen passes over a gas :—

$$2\mathbf{Hg}O = 2Hg + O_2.$$

he chief interest connected with this method is that was the one employed by Priestley: it is seldom opted by chemists, as it is too expensive. The periment here described is more instructive as an cellent case of the resolution by simple means of compound body into its constituents, than valuable a source of oxygen.

2. It may also be prepared by exposing manganic xide (black oxide of manganese), $\mathbf{Mn}O_2$, to a red eat, in an iron bottle or retort :—

$$3\mathbf{Mn}O_2 = \mathbf{Mn}_3O_4 + O_2.$$

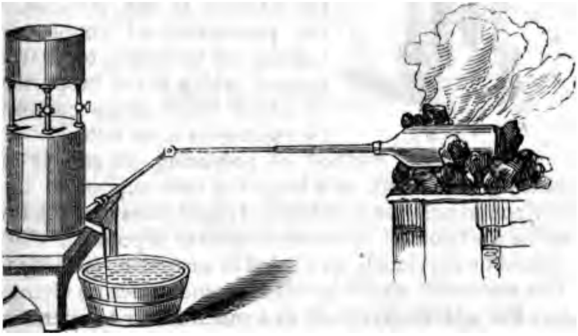
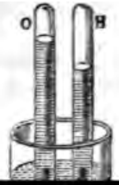


Fig. 4.

In this process one-third of the oxygen is obtained. This method is sometimes adopted when large quantities of oxygen are required, but the oxygen is seldom pure. (See *Fig.* 4.)

3. Oxygen may also be prepared by heating in a glass retort a mixture of manganic oxide and sulphuric acid. These materials should be well mixed, so that no dry place appears in that part of the retort where the materials are placed. If this is not attended to, fracture of the retort is likely to take place. The decomposition is easily illustrated:—

$$2MnO_2 + 2SO_2Ho_2 = 2OH_2 + \underset{\text{Manganous sulphate.}}{2SO_2Mno''} + O_2.$$



4. Oxygen is also prepared by the decomposition of water by voltaic electricity. (See *Fig.* 5.) This is interesting as determining the composition of water, the hydrogen being liberated at the negative electrode, and the oxygen at the positive, in

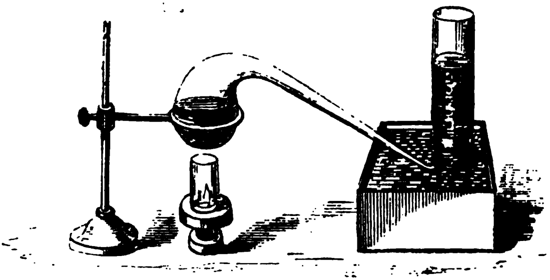


Fig. 6.

paration of gases the first part is a mixture of air, and should always be allowed to escape. Potassic chlorate is composed of 1 atom of potassium, 1 of chlorine, and 3 of oxygen; on the application of heat the whole of the oxygen is given off:—

$$2\left\{\begin{array}{l}OCl\\OKO\end{array}\right. = 2KCl + 3O_2.$$

Potassic chlorate.

6. Oxygen can be obtained by heating a mixture of potassic bichromate and sulphuric acid in a retort or flask:—

$$2\left\{\begin{array}{l}CrO_2Ko\\O\\CrO_2Ko\end{array}\right. + 8SO_2Ho_2 = 2SO_2Ko_2 + \underset{\text{Chromic sulphate.}}{2S_3O_6(Cr_2O_6)^{VI}}$$

$$+ 8OH_2 + 3O_2.$$

7. If a mixture of steam and chlorine be passed through a red-hot porcelain tube, the chlorine unites with the hydrogen of the water, and oxygen is liberated, thus:—

$$2OH_2 + 2Cl_2 = 4HCl + O_2.$$

8. Oxygen can be produced on a large scale by dropping sulphuric acid into a red-hot platinum

or porcelain tube; the sulphuric acid molecule is broken up by the intense heat as under :—

$$2SO_2Ho_2 = 2SO_2 + 2OH_2 + O_2.$$

The sulphurous anhydride is absorbed by suitable means, and the oxygen passed into a gasholder.

Many other processes for making oxygen have from time to time been proposed, but as yet they have no importance.

Oxygen is liberated on the grand scale in nature by the decomposition, through the agency of sunlight, of carbonic anhydride by the green leaves of plants :—

$$CO_2 = C + O_2.$$

Properties.—Oxygen at all observed temperatures is a transparent, colourless, inodorous, and tasteless gas. It is not combustible—that is, it will not burn,—but it is a powerful supporter of the burning of other bodies. It has no acid nor alkaline reaction, although it enters into the composition of all oxygen acids and alkalies. Bodies which burn in air, burn with increased brilliancy in pure oxygen. A strip of wood, if only a spark remain, is instantly rekindled

Fe_3O_4 : the burning phosphorus produces phosphoric anhydride, P_2O_5, in solid white flakes, which dissolve in the water: sulphurous anhydride gas, SO_2, is produced from the sulphur, and carbonic anhydride gas, CO_2, from the burning charcoal. Both the SO_2 and the CO_2 dissolve partly in the water. In all these cases chemical combination takes place between the oxygen and the burning body; and the increase in weight of the body is exactly the same as the weight of oxygen consumed. All the common cases of combustion which take place in air give precisely the same result as those which take place in oxygen. The action in pure oxygen is more energetic, owing to the absence of nitrogen. In the atmosphere these changes take place in a more subdued manner. The respiration of animals is, in effect, the same as common combustion. The substances in the blood burn away slowly by the aid of the oxygen introduced into the system. An ounce of iron rusting in the air produces just the same amount of heat as if it had been burnt like the wire in oxygen. This slow combustion, unattended by light, often extends over several years, while the same result could be produced by pure oxygen (assisted at first by heat) in a few seconds.

The decay of animal and vegetable substances is a process of slow oxidation. The offensive impurities of the air and water are destroyed in a great degree by a slow process of burning. The lungs of land animals, and the gills of fish, are adapted to absorb oxygen, the animal from the air, and the fish from the dissolved oxygen of the water. An animal confined in a given bulk of air dies as soon as the oxygen is consumed; but if confined in the same bulk of oxygen, the circulation quickens, and fever and death are the results,—*not*, however, as in the previous case, from deprivation of oxygen, but from an excess of that body, causing the animal to live *too quickly*. Such is a brief outline of the relation of oxygen to animal life. Its action seems destructive, yet it is the sustainer of life, the mainspring of all vital action. Oxygen may be distinguished from every other gas, except nitrous oxide, by plunging into it a taper which has just been blown out; if the smallest part is incandescent it immediately bursts into flame. Oxygen gives red fumes when mixed with nitric oxide, which nitrous oxide does not.

Ozone—Allotropic Oxygen—O_3.

Molecular weight, 48. Molecular volume, normal. 1 litre weighs 24 criths.

This remarkable gas has been the subject of many researches, especially by Schonbein, of Basle, who gave it the name it bears (from οζω, I smell), on account of its peculiar odour.

Preparation.—1. If a clean piece of phosphorus oil of turpentine, and other essential oils, and Siemens has prepared it by electric induction. The whole of the oxygen cannot, by any known process, be converted into ozone.

Properties.—When heated, ozone is transformed into ordinary oxygen with increase of its volume. It possesses bleaching and disinfecting properties, and hence probably performs useful service in the economy of nature, as it is supposed to exist in the atmosphere. It is an extremely powerful oxidizing agent; it corrodes cork, and rapidly oxidizes iron, copper, and even silver and mercury. It is absorbed by the essences of turpentine and cinnamon. When it oxidizes a substance, it does not diminish in bulk; this is probably to be accounted for by assuming that only one of the atoms in the molecule of ozone, O_3, is employed in oxidizing, and the other two remain to form a molecule of ordinary oxygen, which will occupy the same bulk as the ozone itself: thus,—

$$Ag_2 + \underset{\text{One molecule ozone.}}{O_3} = \mathbf{O}Ag_2 + \underset{\text{One molecule oxygen.}}{O_2}.$$

Test.—A strip of paper moistened with a mixture of starch paste and solution of potassic iodide is turned blue by ozone (see test for iodine):—

$$2KI + \mathbf{O}H_2 + O_3 = 2KHo + I_2 + O_2.$$

The blue substance is a compound of iodine and starch. This test is far from being a satisfactory one, as the paper is turned blue by other substances, especially by certain oxides of nitrogen, which very probably exist in the air in minute quantities: hence the existence of ozone in the air is at the best problematical.

Water—OH_2.

Molecular weight = 18. Molecular volume, normal. 1 litre of water vapour weighs 9 criths. Boils at 100° C.

The composition of water was first discovered by Mr. Cavendish, but its exact composition by weight and volume has occupied the attention of chemists for some years. There are two methods of determining the composition of any compound body: these are *analysis*, by which the compound is resolved into its constituents; and *synthesis*, by which the elements or constituents are made to unite to produce the compound. Both these methods have been brought to bear on the determination of the proportions of oxygen and hydrogen in water.

Analysis.—The most elegant illustration of the analysis of water is its decomposition by voltaic electricity (see p. 66). This experiment proves that water is composed of two volumes of hydrogen united to one volume of oxygen. Methods 1, 2, and 6, for preparing hydrogen, are also processes for determining the composition of water by analysis.

Synthesis.—When a mixture of exactly two volumes of hydrogen and one volume of oxygen is exploded in a strong glass tube standing over mercury, the whole of the mixture becomes water, whereas any proportions varying from these leave an excess of oxygen or hydrogen, as the case may be. The following is another synthetical mode:—Dry hydrogen is passed through a red-hot tube containing dry cupric oxide; the water vapour produced is carefully condensed and weighed; the cupric oxide is also weighed, both before and after the operation: the loss of weight of the cupric oxide is always found to be in proportion to the weight of water contained, as 8 : 9 :—

$$2CuO + 2H_2 = Cu_2 + 2OH_2.$$

In this equation the loss of weight of $2CuO = 32$ and amount of water formed $= 36$; and—

$$\underset{\text{Oxygen.}}{\text{As } 32} : \underset{\text{Water.}}{36} :: \underset{\text{Oxygen.}}{16} : \underset{\text{Water.}}{18};$$

therefore, every sixteen parts by weight of oxygen unite with two parts by weight of hydrogen to form eighteen parts by weight of water.

Properties.—Water, when pure, is colourless, except when viewed in a large mass, when it appears of a beautiful blue colour. It is inodorous, tasteless, and transparent. It attains its maximum density at 4° C.; it boils under ordinary atmospheric pressure at 100°. It evaporates at all temperatures. It is 815 times heavier than air; an imperial gallon weighs 10 lbs. avoirdupois. A cubic metre of distilled water at 4° C. weighs 1,000 kilogrammes, and a litre weighs 1,000 grammes. It is compressible in a very slight degree. The specific gravity of water vapour, when compared with hydrogen at the same temperature and pressure, is $\frac{2 + 16}{2} = 9$.

Water never occurs in nature perfectly pure. Filtration through sand will remove suspended impurities, and distillation will free the liquid from those dissolved. The solvent properties of water are far greater than those of any other liquid. Many salts are soluble in water, the solubility increasing with an increase of temperature, and the solution deposits crystals as it cools. Common salt and some of the salts of calcium are, however, exceptions to this general rule. (See p. 45.)

Water dissolves most of the gases, although in very unequal quantities. Hydrogen, oxygen, and common air are absorbed to a small extent, while ammonia and hydrochloric acid gases are dissolved to a very great

[illegible]

[illegible] is in [illegible] enters into [illegible] is water of [illegible] has already been [illegible]

[illegible] WATERS

[illegible] impu- [illegible] has fallen. [illegible] composition of the [illegible] commonest [illegible] are calcic and magnesic carbo[illegible] sulphates, and sodic and [illegible] compounds, *effervescent* when it evolves gas freely, *sulphureous* when it contains sulphuretted hydrogen in solution, and *saline* when considerable quantities of salts are held in solution, such as magnesic and sodic sulphates.

River water is frequently contaminated with town sewage, manure drainage from land, vegetable matter in solution, and numerous impurities in suspension. It is therefore not so well fitted for domestic purposes as spring water, as organic matter is much more detrimental to health than mineral matter. The *kind* of organic impurity is of more importance than the quantity, as decomposing animal matters are more likely to be prejudicial to health than vegetable substances.

Sea water contains a variety of salts in solution, sodic chloride being the principal one. Magnesic chloride, sulphate, and bromide, and calcic carbonate and sulphate are also present, and traces of iodine and ammonia.

Water for domestic purposes should be bright and clear, and as free as possible from organic impurities. It should not destroy the colour imparted to it by adding a few drops of *Condy's fluid*, nor should it leave much residue when a quart of it is evaporated to dryness in a small porcelain dish; the dry residue should not give off red fumes or the smell of burning hair when heated to redness. The water should have no taste, and should give a lather pretty readily with soap. "We are at present very much in the dark with regard to the sanitary aspect of the amount of solids existing in water. Whether a water having an exceptionally small amount of solid contents is specially salubrious remains to be ascertained. A very large quantity of mixed matter is certainly unwholesome. To take an extreme case, sea water is absolutely nonpotable. But whether the 21 grains per gallon in London water, and whether double that amount, would do the smallest damage to the health of persons who should drink such water, is an open question."* Water is freed from dissolved organic impurities by using animal charcoal for the filter.

Hardness and Softness.—These terms are applied to water according as it produces a lather with difficulty or easily with soap. The hardness arises from the presence of calcic and magnesic salts, which destroy the soap, and it is always estimated in the degrees of Dr. Clark, as given below; the number of grains of calcic carbonate, or its equivalent of other calcic or magnesic salts, giving the number of degrees of hardness: thus Thames water above London contains 15 grains of calcic carbonate per gallon; it is therefore of 15 degrees of hardness.

The degree of hardness is ascertained by means of Dr. Clark's soap test, which is made by dissolving curd soap in rectified spirit, and making it of such a strength that 32 decems (a decem = 10 water grains) five minutes. The degree of hardness is then calculated by Dr. Clark's table as under:—

Degree of hardness.	Soap test decem measures required.	Differences as for the next degree of hardness.
0	1·4	1·8
1	3·2	2·2
2	5·4	2·2
3	7·6	2·0
4	9·6	2·0
5	11·6	2·0
6	13·6	2·0
7	15·6	1·9
8	17·5	1·9
9	19·4	1·9
10	21·3	1·8
11	23·1	1·8
12	24·9	1·8
13	26·7	1·8
14	28·5	1·8
15	30·3	1·7
16	32	

Thus a sample of water, of which 1,000 grains require 15·6 decems of the solution of soap to be added before a lather is formed, is of 7 degrees of hardness. Water which exhibits 7 or more degrees of hardness is said to be *hard* water; if below 7 degrees it is regarded as *soft* water. Water which is over 16 degrees of hardness is too hard for ordinary purposes.

Calcic and magnesic carbonates are almost insoluble in pure water, but they are freely dissolved by carbonic acid, which is usually contained in water. These carbonates give hardness to water, but they may be removed in a great measure by boiling the water, or by adding lime to it: thus,—

$COCao''$	+	$COHo_2$	+	$CaHo_2$	=	$2COCao''$	+	$2OH_2$
Calcic carbonate.		Carbonic acid.		Calcic hydrate.		Calcic carbonate.		Water.

The calcic carbonate already dissolved in the water, together with that formed by the action of the lime, on the dissolved carbonic acid, separate from the water and deposit. As calcic carbonate is therefore removeable, the hardness of the water containing it is said to be *temporary*. Water, however, which is rendered hard by having calcic sulphate in solution, is said to be *permanently* hard, as the calcium salt cannot be removed without introducing other objectionable substances.

Hydroxyl—H_2O_2, or $(HO)_2$, or Ho_2, or $\left\{\begin{matrix}OH\\OH.\end{matrix}\right.$

Hydric peroxide, binoxide of hydrogen, oxygenated water.

This substance has attracted considerable attention recently on theoretical grounds, inasmuch as it is believed to be contained in a vast number of compounds, and to play a very important part in their constitution, and in the chemical changes which they undergo.

all oxy-acids, hydrates, acid salts, and in multitudes of other compounds. For numerous examples of its occurrence the student is referred to the table of salts, p. 33 *et seq.*

Preparation.—Hydroxyl is somewhat difficult to make; the following are the processes usually employed :—

1. Baric peroxide is suspended in ice-cold water, and carbonic anhydride is passed through the water for some time; the containing vessel should be surrounded by ice. The hydroxyl is produced by the following reaction :—

$$\underset{\text{Baric peroxide.}}{\left\{\begin{matrix}O\\O\end{matrix}\right.Ba''} + \underset{\text{Carbonic anhydride.}}{CO_2} + OH_2 = \underset{\text{Baric carbonate.}}{COBao''} + \underset{\text{Hydroxyl.}}{\left\{\begin{matrix}OH\\OH.\end{matrix}\right.}$$

2. Baric peroxide is dissolved in hydrochloric acid, carefully cooled by ice :—

$$\left\{\begin{matrix}O\\O\end{matrix}\right.Ba'' + 2HCl = BaCl_2 + \left\{\begin{matrix}OH\\OH.\end{matrix}\right.$$

The barium is precipitated by cautiously adding sulphuric acid :—

$$BaCl_2 + SO_2Ho_2 = SO_2Bao'' + 2HCl.$$

The chlorine of the hydrochloric acid is then removed by adding solution of argentic sulphate :—

$$SO_2Ago_2 + 2HCl = SO_2Ho_2 + 2AgCl.$$

And, lastly, the sulphuric acid is precipitated by the addition of baryta water :—

$$SO_2Ho_2 + BaHo_2 = SO_2Bao'' + 2OH_2.$$

In all the above stages the insoluble precipitates will require to be filtered off; the last fluid will contain nothing but hydroxyl and water, the latter of which may be removed by placing it under an air-pump, and continuously exhausting until the fluid has acquired the sp. gr. of 1·45.

Properties.—It is a colourless, transparent, odourless, syrupy liquid, possessing powerful bleaching properties. It is very prone to decomposition: it effervesces with the least elevation of temperature, giving off oxygen gas: when raised to 100° it explodes. It is freely soluble in water, and in this state is not so liable to decompose. It has powerful oxidizing properties. It converts plumbic sulphide, PbS, into plumbic sulphate, SO_2Pbo'', which property is turned to account in the restoration of oil paintings which have become blackened by the conversion of the lead into sulphide. Its most remarkable property is its power to reduce argentic oxide to the metallic state:—

$$\left\{\begin{matrix}OH\\OH\end{matrix}\right. + OAg_2 = OH_2 + O_2 + Ag_2.$$

This peculiar reaction gives considerable support to the assumption that elementary bodies, when free, consist of molecules in which two atoms are combined. (See Molecules, p. 8.)

Oxides and Oxy-acids of Chlorine.

These compounds can only be obtained by indirect means.

There are three oxides known,—

Hypochlorous anhydride . .	OCl_2
Chlorous anhydride . . .	$\left\{\begin{matrix}OCl\\O\\OCl\end{matrix}\right.$
Chloric peroxide . . .	$\left\{\begin{matrix}OCl\\O\\O\\OCl\end{matrix}\right.$

And four oxy-acids,—

Hypochlorous acid	$ClHo$
Chlorous acid	$OClHo$
Chloric acid	$\left\{\begin{matrix}OCl\\OHo\end{matrix}\right.$
Perchloric acid	$\left\{\begin{matrix}OCl\\O\\OHo\end{matrix}\right.$

Hypochlorous anhydride, OCl_2, is prepared by passing dry chlorine through a tube containing mercuric oxide, produced by precipitation and subsequent exposure to a strong heat. It is a pale yellow gas, having a peculiar odour distinct from that of chlorine.

$$\underset{\substack{\text{Mercuric}\\\text{oxide.}}}{2HgO} + 2Cl_2 = \underset{\substack{\text{Mercuric}\\\text{oxychloride.}}}{\left\{\begin{matrix}HgCl\\O\\HgCl\end{matrix}\right.} + OCl_2.$$

The gas is decomposed by water, forming *hypochlorous acid*, ClHo, which yields hypochlorites when acted on by metallic oxides or hydrates, thus :—

$$\underset{\substack{\text{Sodic}\\\text{hydrate.}}}{ONaH} + \underset{\substack{\text{Hypoclorous}\\\text{acid.}}}{ClHo} = \underset{\substack{\text{Sodic}\\\text{hypochlorite.}}}{ClNao} + OH_2.$$

The substance called *bleaching powder* was formerly thought to be a mixture of calcic chloride and calcic hypochlorite; it does not appear, however, to contain calcic chloride. It is made by passing chlorine gas into a low room, on the floor of which a layer of lime is placed, until no more gas is absorbed. The reaction seems to be as follows :—

$$\underset{\substack{\text{Calcic}\\\text{hydrate.}}}{CaHo_2} + Cl_2 = \underset{\substack{\text{Calcic}\\\text{chloro-hypochlorite.}}}{Ca(OCl)Cl} + OH_2.$$

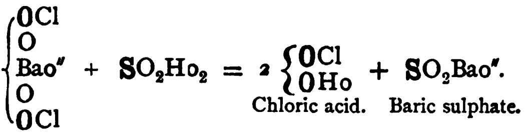
The calcic chloro-hypochlorite is known in commerce under the names of *bleaching powder* and *chloride of lime.* Acids liberate free chlorine from it, thus :—

$$Ca(OCl)Cl + 2HCl = CaCl_2 + OH_2 + Cl_2.$$

Chlorous anhydride, $\left\{\begin{matrix}OCl\\O\\OCl\end{matrix}\right.$, is prepared by heating in a flask a mixture of potassic chlorate, arsenious acid, and nitric acid. The flask must be filled to the neck, and heated in a water bath. The reaction is somewhat complicated. Chlorous anhydride is a greenish yellow gas, which liquefies with extreme cold. It is decomposed by water, forming *chlorous acid,* $OClHo$, a greenish yellow liquid, having strong bleaching and oxidizing properties. With bases it forms chlorites.

Chloric peroxide, $\left\{\begin{matrix}OCl\\O\\O\\OCl\end{matrix}\right.$, is made by very cautiously heating in a small glass retort a mixture of powdered potassic chlorate and sulphuric acid made into a paste. The retort is placed in warm water.

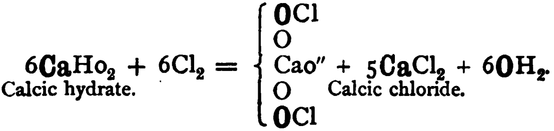
Chloric acid, $\left\{\begin{matrix}OCl\\OHo\end{matrix}\right.$, is obtained by adding dilute sulphuric acid to a solution of baric chlorate so long as a precipitate separates. The solution contains chloric acid : it must not be filtered, but allowed to settle, and the liquid decanted from the insoluble baric sulphate :—



$$\underset{\text{Baric chlorate.}}{\left\{\begin{matrix}OCl\\O\\Bao''\\O\\OCl\end{matrix}\right.} + SO_2Ho_2 = 2\underset{\text{Chloric acid.}}{\left\{\begin{matrix}OCl\\OHo\end{matrix}\right.} + \underset{\text{Baric sulphate.}}{SO_2Bao''.}$$

The liquid may be concentrated by cautious evaporation. It is then a syrupy fluid, having extremely powerful oxidizing properties. It is readily decomposed.

Chlorates may be made by passing chlorine gas through a solution of a hydrate : thus calcic chlorate is made by passing chlorine through boiling milk of lime :—



$$\underset{\text{Calcic hydrate.}}{6CaHo_2} + 6Cl_2 = \underset{\text{Calcic chlorate.}}{\left\{\begin{matrix}OCl\\O\\Cao''\\O\\OCl\end{matrix}\right.} + \underset{\text{Calcic chloride.}}{5CaCl_2} + 6OH_2.$$

Potassic chlorate is obtained by the addition of potassic chloride to the above solution of calcic chlorate. The potassic chlorate is separated from the calcic chloride by crystallization.

Perchloric acid, $\left\{\begin{matrix}\text{OCl}\\ \text{O}\\ \text{OHo}\end{matrix}\right.$, is prepared by distilling potassic perchlorate with sulphuric acid:—

$$2\left\{\begin{matrix}\text{OCl}\\ \text{O}\\ \text{OKo}\end{matrix}\right. + \text{SO}_2\text{Ho}_2 = 2\left\{\begin{matrix}\text{OCl}\\ \text{O}\\ \text{OHo}\end{matrix}\right. + \text{SO}_2\text{Ko}_2.$$

Potassic perchlorate. Perchloric acid.

Pure perchloric acid is a colourless liquid; it soon becomes coloured, and after a few weeks decomposes with explosion. When brought in contact with charcoal or other organic substances it explodes with fearful violence. It should therefore only be made in small quantities, and then cautiously.

Potassic perchlorate is made by heating potassic chlorate until one-third of the oxygen has come off; the substance then consists of a mixture of potassic chloride and perchlorate, which may be separated by crystallization:—

$$2\left\{\text{OCl}\right. = \text{KCl} + \left\{\begin{matrix}\text{OCl}\\ \text{O}\end{matrix}\right. + \text{O}$$

Preparation.—Like oxygen, boron is capable of assuming more forms than one; it can be obtained in three allotropic forms:—

Amorphous boron is obtained by heating boric anhydride with sodium:—

$$B_2O_3 + 3Na_2 = 3ONa_2 + B_2.$$

It is a brownish powder; heated in air it burns to boric anhydride.

Graphitoid boron is got when boric chloride is passed over fused aluminium:—

$$Al_2 + 2BCl_3 = Al_2Cl_6 + B_2.$$

It crystallizes in thin opaque six-sided plates.

Diamond boron is made by fusing boric anhydride with aluminium:—

$$Al_2 + B_2O_3 = Al_2O_3 + B_2.$$

The crystals are in square octahedrons, of a brownish colour, and exceedingly hard, as they will scratch the ruby. Sp. gr., 2·68.

Boric anhydride, B_2O_3, is obtained by fusing boric acid at a red heat:—

$$2BHo_3 = B_2O_3 + 3OH_2.$$

The anhydride is a glassy-looking substance.

Boric acid, BHo_3 (boracic acid), is most readily obtained by adding hydrochloric acid to a hot solution of *borax:* as the liquid cools the boric acid separates in beautiful crystals. When cold the liquid may be decanted, and the crystals slightly washed with cold water:—

$$\underset{\text{Borax.}}{B_4O_5Nao_2} + 2HCl + 5OH_2 = \underset{\text{Boric acid.}}{4BHo_3} + 2NaCl.$$

On raising a solution of boric acid to 100° C. it is converted into metaboric acid, **B**OHo, with loss of a molecule of water.

Boric chloride, $\mathbf{B}Cl_3$, is prepared by exposing to the action of dry chlorine at a very high temperature an intimate mixture of fused boric anhydride and charcoal:—

$$\underset{\text{Boric anhydride.}}{\mathbf{B}_2O_3} + 3Cl_2 + C_3 = \underset{\text{Boric chloride.}}{2\mathbf{B}Cl_3} + \underset{\text{Carbonic oxide.}}{3\mathbf{C}''O}.$$

Borax, $B_4O_5Nao_2$ is an abnormal sodic borate; it is employed for soldering metals, forming the glazes for porcelain, and as a flux for examining metallic compounds before the blowpipe. A loop made of platinum wire is strongly heated and dipped into borax powder; a series of experiments can now be made. A salt of manganese heated in the outer flame of the blowpipe, with the borax on the wire, gives a violet-coloured bead; cobalt, blue; chromium, green.

Boron, when heated, will burn in nitrogen gas, forming therewith boric nitride, **B**N.

TETRAD ELEMENTS.

Carbon—C.

of all animal and vegetable substances. *Coal* consists mainly of carbon,—and some varieties, as *anthracite*, almost entirely of it.

Preparation and Properties.—Charcoal is obtained by exposing animal or vegetable substances to the action of heat, carefully excluding the atmosphere during the operation. This process is now usually performed in large iron retorts, somewhat similar to those which are used for making *coal gas*. The specific gravity of the charcoal varies with the wood from which it is manufactured.

Charcoal is very combustible, and is extensively used on the Continent for fuel. When perfectly pure it burns without flame. It is a very indestructible and unchangeable substance. Wooden piles driven into the earth are usually charred to prevent decay. Charcoal has the power of condensing in its pores different gases. It has been found to absorb 90 times its bulk of ammonia, 55 of sulphuretted hydrogen, 35 of carbonic anhydride, and 9¼ of oxygen. The absorbed gases are condensed within the pores of the charcoal, and those which are combustible are burnt up by the oxygen which is also present. In this manner charcoal acts as an oxidizer, and it is in virtue of this property that it is a deodorizer. Charcoal when newly prepared restores meat slightly tainted, either by being boiled with it or rubbed over it. It thus removes all offensive odours. Rats buried in powdered charcoal yielded no unpleasant effluvia when opened after three months. Water is purified by passing it through filters constructed of layers of charcoal and sand. Black crayons are made from the charcoal of the willow. The charcoal used in the manufacture of gunpowder is made from dogwood or elder.

Ivory and bone-black are varieties of charcoal which result in the heating of bones in a retort till all the volatile products are evolved. This bone-black, however, only contains about 10 per cent. of carbon. The remainder is chiefly calcic phosphate and carbonate. Ivory-black is used as ink in printing from copper and steel engravings. The preparation of bone-black is now an important branch of chemical manufacture, as it is largely used in sugar-refining, owing to its bleaching or decolourizing properties.

Lamp-black results from the imperfect combustion of tar, oil, or resin. The greater part of the carbon is carried off in a state of finely divided powder. When these particles come in contact with a cold body they form a deposit of carbon, which may be scraped off. This forms the lamp-black of commerce. The operation is performed with a limited supply of air in a high chimney, containing a number of plates, upon which the lamp-black collects like soot in our own chimneys.

Plumbago or *black-lead* was formerly thought to be a carbide of iron; it is, however, little else but

For diamonds of moderate size, and the same quality, the value is estimated in this manner:—The weight is given in carats. These carats are subdivided into halves, quarters, eighths, sixteenths, and thirty-second parts. If four diamonds weigh 1, 2, 3, and 4 carats each, their respective value will be as 1, 4, 9, and 16; their value is in the ratio of the squares of their respective weights. The Koh-i-Noor weighed 186 carats; since it has been recut it weighs $102\frac{13}{16}$ carats.

Carbon, like boron, assumes three allotropic forms; in the artificial varieties, such as charcoal and lamp-black, we have *amorphous* carbon. *Black-lead*, or *graphite*, as it is sometimes called, is a crystallized form of carbon, the crystals being six-sided plates belonging to the rhombohedral system. *Gas carbon*, a very hard form of this variety of carbon, is deposited in the hottest part of gas retorts, and results from the partial decomposition of the coal gas when the retorts are made too hot; it forms the carbon element in Bunsen's form of galvanic battery. Graphite is also one of the constituents of cast iron. The other crystallized form of carbon is the *diamond*, the crystals of which belong to the regular system; carbon is therefore *dimorphous*,—that is, it assumes two crystalline forms.

COMPOUNDS OF CARBON AND OXYGEN.

There are two well-known compounds of oxygen and carbon—carbonic anhydride and carbonic oxide.

Carbonic anhydride, CO_2, is an important constituent of the atmosphere, of which it forms $\frac{1}{2500}$ of its bulk in the open country; it is also present in all spring water; it is evolved from volcanoes, and is a constant product of respiration. It is the dreaded *choke-damp* of the miner. Fermenting liquids, decaying animal and vegetable substances, and the combustion of coal, wood, and gas, are continual sources of carbonic anhydride.

The acid corresponding to CO_2 has the composition $COHo_2$, but it has never yet been isolated, on account of its decomposition into OH_2 and CO_2. Its salts, however, are well known, and are called carbonates; thus we have the two sodic carbonates, $CONao_2$ and $COHoNao$, and calcic carbonate, $COCao''$.

Preparation.—Carbonic anhydride may be prepared *synthetically* by burning carbon in oxygen, or *analytically* by decomposing a carbonate. When carbon is burnt in oxygen no change of volume takes place, although there is an increase of weight; for one volume of hypothetical carbon vapour unites with *two* volumes of oxygen, yielding *two* volumes of carbonic anhydride: the sp. gr. of the gas will therefore be $\frac{12 + (16 \times 2)}{2} = 22$; that is, one volume of this gas weighs 22 times as much as the same volume of hydrogen.

a jar of carbonic anhydride. It is not a supporter of combustion or respiration, nor is it combustible. Heated potassium or sodium, or previously ignited magnesium, will, however, burn in carbonic anhydride, the carbon being deposited as a black mass:—

$$3CO_2 + 2K_2 = 2COKo_2 + C.$$

It extinguishes flame,—not for the same reason as nitrogen, by excluding oxygen; for if we mix one volume of oxygen with four of carbonic anhydride, it does not support combustion, although the quantity of oxygen is as great as in common air. If a living animal be immersed in an atmosphere of nitrogen, it dies the same as if it were immersed in water. The nitrogen kills by suffocation; carbonic anhydride kills by poisoning: all the other conditions of life may be present, but the inhalation of this gas produces death. In winter, on the Continent, people burn charcoal in braziers,—a most unhealthy and dangerous practice where there is not a proper exit for the carbonic anhydride produced.

A solution of carbonic anhydride reddens litmus. Carbonic acid is a dibasic acid, giving rise to two classes of salts, in one of which half the hydrogen of the acid, and in the other the whole of it, is replaced by a metal or radical; thus we have hydric-potassic carbonate, $COHoKo$, and potassic carbonate, $COKo_2$, the original acid being $COHo_2$.

Tests.—Carbonic anhydride is expelled from carbonates on the addition of a strong acid: the gas passed into calcic hydrate (lime water) gives a white precipitate of calcic carbonate (*chalk*), which re-dissolves with an excess of the gas:—

$$CaHo_2 + CO_2 = COCao'' + OH_2.$$

The same effect is produced when the breath is blown through lime water by means of a glass tube. The carbonic anhydride in the breath unites with the lime.

Carbonic Oxide, $C''O$.—The blue lambent flame sometimes seen on the top of brick-kilns and coke-ovens, and in our own fireplaces, is due to the combustion of carbonic oxide. This compound contains one atom less of oxygen than carbonic anhydride, and the carbon is consequently here a diad.

Preparation.—1. It may be prepared by passing carbonic anhydride through a red-hot iron tube filled with iron borings:—

$$4CO_2 + Fe_3 = Fe_3O_4 + 4CO.$$

$4CO_2$	Fe_3	Fe_3O_4	$4CO$
Carbonic anhydride.		Magnetic oxide of iron.	Carbonic oxide.

2. It may also be obtained by passing carbonic anhydride through a red-hot porcelain tube filled with charcoal:—

$$CO_2 + C = 2CO.$$

CO_2	C	$2CO$
Carbonic anhydride.		Carbonic oxide.

A similar change to this takes place in charcoal and [redacted] this case, at a boiling temperature, it withdraws a molecule of water from the oxalic acid, leaving C_2O_3, which cannot exist as such, but breaks up into a mixture of carbonic oxide, CO, and carbonic anhydride, CO_2. By passing the gases through a solution of caustic potash, the carbonic anhydride is absorbed, and the carbonic oxide, CO, passes over, and may be collected at the pneumatic trough in the usual way.

$$\underset{\text{Oxalic acid.}}{\left\{\begin{matrix}COHo\\COHo\end{matrix}\right.} = OH_2 + \underset{\text{Carbonic oxide.}}{C'O} + \underset{\text{Carbonic anhydride.}}{C^{IV}O_2}.$$

Carbonic oxide may be obtained by many other processes; thus it is formed when carbon burns in a limited supply of air, when formic acid or a formiate is heated with sulphuric acid, when potassic ferrocyanide is heated with sulphuric acid, or by heating iron, zinc, or carbon with a carbonate.

Properties.—This gas is transparent, colourless, tasteless, and inodorous; it is slightly lighter than common air. Its sp. gr. is $\frac{12 + 16}{2} = 14$. It is highly combustible, but does not support either combustion or respiration. It burns, in air or oxygen, with a blue lambent flame, becoming carbonic anhydride. Dumas has shown that this gas is a hundred times more poisonous than carbonic anhydride, a fact which till lately was unknown. It acts as a reducing agent at high temperatures in many metallurgical operations.

The compounds of carbon with hydrogen, chlorine, and nitrogen, belong to organic chemistry.

PENTAD ELEMENTS.

Nitrogen—N_2.

Antomic weight and sp. gr. = 14. Molecular weight = 28. Molecular volume, normal. 1 litre weighs 14 criths. Atomicity, V. (frequently III., sometimes I.).

Occurrence in Nature.—Nitrogen forms by volume about 4-5ths of the atmosphere, and enters into a variety of combinations. It exists in combination in the vegetable kingdom, forming either nutritious food like gluten, useful medicines like quinine, or violent poisons like prussic acid, strychnia, and nicotine. It is a constituent of muscular tissue to the extent of about 17 per cent. In the mineral kingdom it is contained in coal, and in potassic and sodic nitrates.

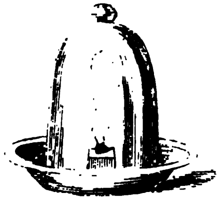


FIG. 8

Nitrogen may be prepared,—of (1.) By burning a small piece of phosphorus in a confined portion of common air under a gas jar standing over water, by which phosphoric anhydride is produced, but the nitrogen remains unacted on. The white solid flakes of phosphoric anhydride are dissolved in the water over which the jar is placed. (See *Fig.* 8.) It must stand till it becomes clear. The water will rise up in the jar to the extent of about 1-5th, and will thus replace the oxygen which is burnt out.

(2.) By passing chlorine through a solution of ammonia:—

$$8NH_3 + 3Cl_2 = 6AmCl + N_2.$$

The HCl thus formed unites with the excess of ammonia, forming NH_4Cl (ammonic chloride). Great care must be taken in this experiment to keep a large excess of ammonia, or the excess of chlorine will act on the ammonic chloride thus:—

$$\mathbf{N}^{v}H_4Cl + 6Cl = 4HCl + \underset{\text{Nitrous chloride.}}{\mathbf{N}'''Cl_3}.$$

This nitrous chloride is a terribly explosive and dangerous compound.

(3.) By mixing iron filings and sulphur with water, and allowing the mixture to remain under a gas jar for twenty-four hours. The oxygen of the atmosphere is absorbed, and nitrogen only remains.

(4.) Nitrogen can also be obtained from the atmosphere by passing air through a hard glass tube containing finely divided copper heated to redness. The copper takes up the oxygen, but the nitrogen passes forward and may be collected:—

$$2Cu + \underbrace{4N_2 + O_2}_{\text{Air.}} = 2CuO + 4N_2.$$

(5.) By heating ammonic nitrite:—

$$\underset{\text{Ammonic nitrite.}}{\mathbf{N}''O(N^{v}H_4O)'} = N_2 + 2OH_2.$$

Properties.—Nitrogen is a permanent gas, having neither colour, taste, nor smell. It is slightly lighter than air. It is not poisonous, but yet will not support respiration. It will not burn unless under exceptional circumstances, nor yet will it support combustion. It is a remarkably inactive body, but many of its compounds have very active and striking qualities. It is estimated that six tons of air pass through a *blast-furnace* every hour. While the oxygen is most active in forming new compounds the nitrogen escapes unchanged. Its chief function in the air is to modify the violent action of the oxygen, and render the atmosphere suitable to the wants of life.

The atmosphere, which is the great storehouse of nitrogen, is composed of a mixture of nitrogen and oxygen, with variable proportions of carbonic anhydride, aqueous vapour, and traces of ammonia, &c.

The analysis of air may be effected in the following manner:—Introduce, by means of a wire, into a graduated tube standing over water, a piece of phosphorus. The known quantity of air can be determined by observing the water point in the graduated tube. The arrangement should now be left for twenty-four hours. The whole of the oxygen will be absorbed by slow combination with the phosphorus. The remaining gas may now be read off, which is nitrogen. Another plan is to mix air with excess of hydrogen in a graduated tube called a *eudiometer*, and explode by an electric spark; the hydrogen and oxygen combine, forming water, which condenses, and the remaining volume of gases is read off: the loss of bulk is due to the whole of the oxygen, *plus* twice as much hydrogen; and therefore the actual amount of oxygen can be ascertained by taking one-third of the loss experienced by the mixed gases during the explosion. (*See paragraph on the composition of gases.*) The amount of oxygen may also be ascertained by the process given under the fourth mode of making nitrogen. The amount of moisture is ascertained by instruments called *hygrometers*, and the carbonic anhydride by passing a known quantity of dry air through a bulb tube containing potassic hydrate, the tube being weighed both before and after the operation. The increase in weight gives the carbonic anhydride.

The following is the average per-centage composition of the atmosphere by volume:*—

---

* Miller's Chemistry, vol. ii.

Nitrogen		77·95
Oxygen		20·61
Aqueous vapour (variable)		1·4
Carbonic anhydride		·04
Nitric acid Ammonia Marsh gas		traces
Sulphuretted hydrogen Sulphurous anhydride	near towns	traces
		100·00

When freed from moisture and carbonic anhydride, the following is its per-centage composition :—

	By weight.	By volume.
Nitrogen	76·9	79·1
Oxygen	23·1	20·9
	100	100

We have already seen that animals withdraw oxygen from the atmosphere and pass into it carbonic anhydride ; on the other hand, plants remove carbonic anhydride from it by means of their leaves, and exhale oxygen. Thus the composition of the atmosphere is beautifully balanced, and it never varies in the open country.

That the atmosphere is a mixture and not a compound is proved (1) by the fact that its density is the same as a mixture of the two gases in the same proportions as they are contained in the atmosphere, and (2) the mixture has all the properties of atmospheric air.

Oxides and Oxy-acids of Nitrogen.

We are acquainted with five distinct compounds of nitrogen with oxygen; two of them, being anhydrides, form acids by combining with water:—

Nitrous oxide . . .	ON_2
Nitric oxide . . .	$\begin{cases} NO \\ NO \end{cases}$
Nitrous anhydride . .	$\begin{cases} NO \\ O \\ NO \end{cases}$
Nitric peroxide . .	$\begin{cases} NO_2 \\ NO_2 \end{cases}$
Nitric anhydride	$\begin{cases} NO_2 \\ O \\ NO_2 \end{cases}$
Nitrous acid . . .	$NOHo$
Nitric acid . . .	NO_2Ho

These compounds are a good illustration of what is called the law of *chemical combination in multiple proportions.* It will be seen that the oxygen contained in them is in the proportion of the numbers 1, 2, 3, 4, and 5, to one and the same quantity when inspired. It is prepared by heating moderately in a retort ammonic nitrate: 4 oz. of this salt produce 1 cubic foot of gas. It should be collected over warm water, or over water saturated with the gas, and care should be taken that the salt is pure when the gas made from it is intended to be breathed.

$$\mathbf{N}^{\mathrm{v}}O_2(N^{\mathrm{v}}H_4O) = 2\mathbf{O}H_2 + \mathbf{O}N_2.$$

Nitrous oxide is a transparent and colourless gas, having a faint sweet taste. It supports combustion like oxygen, but it does not produce red fumes when mixed with nitric oxide, and it is much more soluble than oxygen in cold water. It is used by dentists and surgeons for producing insensibility to pain for a short time. Such substances are called anæsthetics.

Nitric Oxide, $\left\{\begin{matrix}\mathbf{N}O\\ \mathbf{N}O\end{matrix}\right.$ —The molecular volume of this gas does not accord with the *law of Ampère*, for its molecule measures 4 volumes. It is probably a case of dissociation at a very low temperature. It may be prepared by introducing into a hydrogen apparatus some copper turnings, a little water, and nitric acid, added by the funnel. This gas may be collected over water. This reaction may best be viewed in this way:—

(1.) Three atoms of copper replace six atoms of hydrogen, thus:—

$$\underset{\text{Nitric acid.}}{6\mathbf{N}O_2Ho} + 3Cu = \underset{\text{Cupric nitrate.}}{3\left\{\begin{matrix}\mathbf{N}O_2\\ Cuo''\\ \mathbf{N}O_2\end{matrix}\right.} + 3H_2.$$

(2.) The nascent hydrogen acts on nitric acid thus:—

$$3H_2 + \underset{\text{Nitric acid.}}{2\mathbf{N}O_2Ho} = 4\mathbf{O}H_2 + \underset{\text{Nitric oxide.}}{\left\{\begin{matrix}\mathbf{N}O\\ \mathbf{N}O\end{matrix}\right.}.$$

The complete change may be expressed in one equation, thus :—

$$8\mathbf{N}O_2Ho + 3Cu = 3\left\{\begin{matrix}\mathbf{N}O_2\\ Cuo''\\ \mathbf{N}O_2\end{matrix}\right. + \left\{\begin{matrix}\mathbf{N}O\\ \mathbf{N}O\end{matrix}\right. + 4OH_2$$

The gas obtained in this manner is colourless and transparent ; but in contact with air or oxygen it produces deep red fumes, which are sufficient to distinguish it from all other gases. A very pretty experiment may be made thus :—Nearly fill a glass cylinder with nitric oxide over water, and by its side fill another smaller cylinder with oxygen. Then allow a few bubbles of oxygen to pass into the cylinder containing the nitric oxide. Deep red fumes immediately appear from the formation of nitrous anhydride and nitric peroxide. These, however, gradually disappear, being absorbed by the water. Pass a few more bubbles, and the same phenomenon will be repeated, till at last the whole cylinder will be filled with water. This gas performs the important part of a carrier of oxygen in the manufacture of sulphuric acid.

A mixture of nitric oxide and carbonic disulphide vapour burns with a brilliant white light.

dride. Liquid nitrous anhydride boils at 0° C., emitting orange-red fumes.

Nitrous Peroxide, $\left\{\begin{matrix}\mathbf{N}O_2\\ \mathbf{N}O_2\end{matrix}\right.$, called also *nitrous gas*, has a normal molecular volume at 0° C., but it is anomalous (4 volumes) at 100° C.: the gas evidently suffers dissociation between these temperatures, and it may be fairly assumed that nitric oxide dissociates at a still lower temperature, otherwise this body will be an exception to important laws which are apparently absolute.

It is prepared by heating plumbic nitrate in a hard glass retort:—

$$\underset{\text{Plumbic nitrate.}}{2\mathbf{N}_2O_4Pbo''} = 2\mathbf{Pb}O + \underset{\text{Nitric peroxide.}}{2\mathbf{N}_2O_4} + O_2.$$

Nitric peroxide at ordinary temperatures is a deep brownish red gas, having a suffocating odour. It is soluble in water, and converts metallic hydrates and oxides into a mixture of nitrates and nitrites. It has neither acid nor basic properties.

Nitric Anhydride, $\left\{\begin{matrix}\mathbf{N}O_2\\ O\\ \mathbf{N}O_2\end{matrix}\right.$, is prepared by passing dry chlorine gas over dried argentic nitrate heated to 95° C.; the anhydride is condensed by a freezing mixture:—

$$\underset{\substack{\text{Argentic}\\ \text{nitrate.}}}{4\mathbf{N}O_2Ago} + 2Cl_2 = 4AgCl + \underset{\substack{\text{Nitric}\\ \text{anhydride.}}}{2N^{v}_2O_5} + O_2.$$

It is a very unstable white crystalline substance.

Nitrous Acid, $\mathbf{N}OHo$, results from the action of water on nitrous anhydride. It forms nitrites with oxides and hydrates.

Potassic nitrite is made by heating nitre to redness.

Nitrites are frequently found in water which has been contaminated with sewage, &c.

Nitric Acid, NO_2Ho, known as *aquafortis*, is

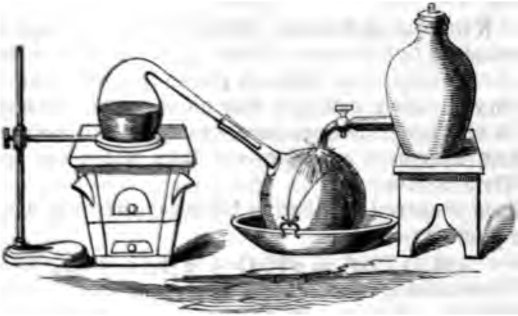


FIG. 9.

one of the most important compounds of nitrogen. In hot dry climates the soil is sometimes covered with a white, downy-looking salt. When this salt is collected, dissolved, filtered, and crystallized, it produces the *nitre* or *saltpetre* of commerce: it is potassic nitrate. To obtain nitric acid from this salt, mix it with half its weight of sulphuric acid in a glass retort (*Fig.* 9): cover the receiver (which may be a Florence part of the nitric acid, and in practice is very liable to injure the retorts. It is therefore customary to use only half this quantity of nitre, when the decomposition is as follows:—

$$\underset{\text{Potassic nitrate.}}{NO_2Ko} + SO_2Ho_2 = \underset{\text{Nitric acid.}}{NO_2Ho} + \underset{\text{Hydric-potassic sulphate.}}{SO_2HoKo.}$$

In this case only half the hydrogen of the sulphuric acid is replaced, and the temperature required is much lower.

On a large scale iron retorts and earthenware condensing vessels are used, and sodic nitrate substituted for potassic nitrate. The iron retort is protected from the action of the acid by lining it with thin fire-bricks.

Pure nitric acid is a colourless liquid. It fumes when exposed to air. It is partially decomposed by light. It is intensely acid, and reddens vegetable blues. It stains the skin and nails yellow. It is used for assaying, and as a solvent for tin in calico-printing. It is a powerful oxidizing agent, attacking most of the metals with great violence.

It forms nitrates with bases, nearly all of which are soluble in water. Mixed with hydrochloric acid it forms *aqua regia*. All the metals except gold and plantinum, and a few rare ones, attack nitric acid. The metals have a different action upon nitric acid from what they have on such acids as sulphuric and hydrochloric. The metals no doubt displace hydrogen from one portion of the nitric acid, but the nascent hydrogen reduces or removes oxygen from another portion. Different metals act very differently on nitric acid, for although the nitrate of the metal is usually produced, yet the secondary products of nitrogen vary: thus, nitric acid yields nitrous oxide, nitric oxide, nitrous anhydride, and nitric peroxide, when acted on by zinc, copper or mercury, silver, and tin respectively.

Nitric acid produces substitution compounds of great utility, such as *gun-cotton*, *nitro-glycerine*, *nitro-benzol*, and *picric acid*. It is the most important oxidizing agent the chemist possesses. It converts ferrous into ferric salts, stannous into stannic, and arsenious into arsenic salts.

Tests.—1. Nitric acid, or a solution of a nitrate mixed with sulphuric acid, turns brown when a strong solution of ferrous sulphate is mixed with it.

2. Copper filings and strong sulphuric acid heated in a test-tube with a nitrate, liberate nitric oxide, recognised by forming red fumes of nitric peroxide in the upper part of the tube.

3. Solution of indigo boiled with HCl and a nitrate is decolourized.

Compounds of Nitrogen with Hydrogen.

Ammonia, $N'''H_3$.—This is the only compound of nitrogen and hydrogen which has been isolated. It is one of the products of the decomposition of animal matters and vegetables containing nitrogen, and is therefore found in the atmosphere; it is brought down by rain, and plays a very important part in the nourishment of plants, which derive most of their nitrogen from this source.

Preparation.—1. Ammonic chloride (sal ammoniac), NH_4Cl, is heated in a flask with quicklime, and the evolved gas is collected over mercury in a small pneumatic trough, or if a solution be required it is passed directly into water, in which it is very soluble:—

$$\begin{array}{ccccccccc} 2NH_4Cl & + & CaO & = & CaCl_2 & + & 2NH_3 & + & OH_2. \\ \text{Ammonic chloride.} & & \text{Calcic oxide.} & & \text{Calcic chloride.} & & \text{Ammonia.} & & \end{array}$$

The principal part of the ammonia now manufactured is obtained from *gas liquor*, a by-product in the manufacture of coal gas, containing carbonate, sulphide, and other ammoniacal salts. The liquor is mixed with lime, and distilled in large iron retorts or boilers: the distillate is very impure, and requires to be distilled a second time with lime, the gaseous product being passed through a lime box purifier, and then into water, which is thus converted into ammonic hydrate (*liquid ammonia*), NH_4Ho ($=AmHo$).

The gaseous products of the first distillation are converted into ammonic sulphate, a very useful salt, by passing them directly into sulphuric acid. The sulphate crystallizes, and is fished out from time to time with perforated shovels.

Ammonic sulphate is largely used as a manure, also in the manufacture of ammonic carbonate and chloride, and also as a cheap substitute for ammonic chloride in the manufacture of ammonic hydrate:—

$$\begin{array}{ccccccc} SO_2Amo_2 & + & CaHo_2 & = & SO_2Cao'' & + & 2AmHo. \\ \text{Ammonic sulphate.} & & \text{Slaked. lime.} & & \text{Calcic sulphate.} & & \text{Ammonic hydrate.} \end{array}$$

2. Organic compounds containing nitrogen, such as bones and animal refuse, when heated with alkaline hydrates, or *soda lime*, yield all their nitrogen as ammonia. When submitted to destructive distillation in a closed vessel they yield ammonia, and some ammonic salts.

Properties.—The gas has a molecular weight of 17, and its molecular volume is normal. 1 litre weighs 8·5 criths. It liquefies under great pressure, the liquid producing great cold by its evaporation, as it boils at —38·5° C. It is a colourless and transparent gas, having an extremely powerful and characteristic odour. At 0° C. water dissolves 1,050 times its own volume of the gas, and at 25° C. 580 times its volume. The ammonic hydrate thus produced is lighter than water: the strongest sold in the shops has a sp. gr. of ·880.

Ammonia was called by the alchemists spirits of hartshorn, because it was obtained by the distillation of deers' horns in close vessels. It is sometimes called the volatile alkali.

Ammonia is a powerful base: it combines with and neutralizes the strongest acids, forming thereby compounds known as ammonic salts, thus:—

$$N'''H_3 + HCl = N^{v}H_4Cl\ (=AmCl).$$

Ammonic chloride.

$$2NH_3 + SO\ Ho_2 = SO_2Amo_2.$$

Ammonic sulphate.

$$NH_3 + NO_2Ho = NO_2Amo.$$

Ammonic

and sodium. It forms an amalgam with mercury, having metallic lustre, and all the physical characters of any other amalgam; on this account it is regarded as a metal by some chemists.

Ammonium amalgam is most readily obtained by warming a small quantity of mercury in a test-tube, and dropping in small bits of metallic sodium. The two metals will coalesce with great evolution of heat and a vivid light. The sodium amalgam, while still hot and fluid, should be poured on a stone or iron slab: it will soon solidify, and if a piece of it be put into solution of ammonic chloride the following reaction will take place:—

$$\underset{\substack{\text{Sodium}\\ \text{amalgam.}}}{Hg_xNa_y} + yNH_4Cl = \underset{\substack{\text{Ammonium}\\ \text{amalgam.}}}{Hg_x(NH_4)_y} + yNaCl.$$

The ammonium amalgam thus formed swells out enormously, but readily breaks up into ammonia, hydrogen, and mercury.

Ammonium forms the compound basylous element in all ammonic salts.

Ammonic Chloride (Sal Ammoniac), AmCl or NH_4Cl, is prepared from gas or bone liquor. The solution is treated with a slight excess of hydrochloric acid, HCl, by which the ammonic carbonate, $COAmO_2$, and ammonic sulphide, SAm_2, always present in these liquids, are decomposed, with evolution of carbonic anhydride, CO_2, and sulphuretted hydrogen, SH_2. The liquid is evaporated to dryness, and the salt carefully heated, to expel the tarry matter with which it is associated. It is then purified by sublimation in iron vessels lined with clay and covered with domes of lead. This salt is sublimed without decomposition. Ammonic chloride forms double salts with the chlorides of magnesium, nickel, cobalt, zinc, copper, manganese, and platinum; forming ($MgCl_2,2NH_4Cl$); ($NiCl_2,2NH_4Cl$); ($CoCl_2,2NH_4Cl$); &c.; ($PtCl_4,2NH_4Cl$). The sal-ammoniacum of the Romans was ammonic chloride, and was obtained as a natural product from a district in Libya.

Ammonic Sulphate, SO_2Amo_2, or $SO_2(NH_4O)'_2$. —The manufacture of this important article on the large scale has already been described (p. 105). A purer substance may be obtained by neutralizing ammonic carbonate, $COAmo_2$, with sulphuric acid. The salt is soluble in two parts of cold water, and crystallizes in long six-sided prisms, which lose an atom of water when heated. Ammonic sulphate is entirely decomposed by ignition; and by long boiling with water, ammonia is volatilized, and the liquid rendered acid.

Ammonic Carbonates.—Deville states that there exist only two of these carbonates, although it has long been supposed that there are a considerable number. 1. Hydric-ammonic carbonate, $COHoAmo$ (bicarbonate of ammonia). This salt is obtained by saturating ammonic hydrate with carbonic anhydride. The sesquicarbonate when exposed to the air gradually loses carbonic anhydride, and is transformed into this salt. 2. Ammonic carbonate (sal volatile), $COAmo_2$, has not been obtained in the solid state.

Commercial ammonic carbonate, called sesquicarbonate of ammonia, contains the elements of two molecules of (1) and one molecule of (2). It is therefore a mixture of the two kinds ($2COHoAmo + COAmo_2$.). It is prepared by heating to redness in a retort a mixture of two parts of ammonic chloride and one part of chalk. The salt sublimes, and is condensed in a receiver which is luted to the retort.

Ammonic Nitrate, NO_2Amo, is prepared by adding dilute [illegible] solution is neutralized [illegible] moderate temperature [illegible] six-sided prisms [illegible] like that salt. When heated [illegible] nitrous oxide and water [illegible]

When ammonia [illegible] ized by hydro[illegible] acid [illegible] SH_2; hydro[illegible] acid [illegible] HBr; the resulting [illegible] ammonium NH_4 [illegible] respective acid [illegible]

Hydro-ammonium sulphide [illegible] pound of great [illegible] saturating a solution [illegible] sulphuretted hydrogen [illegible] many metals from their [illegible] cipitates having characteristic [illegible] to light it becomes [illegible] following decomposition [illegible]

$$4SH_4 [illegible]$$

[illegible]

The yellow [illegible] the laboratory [illegible] half the sulphur [illegible] is evolved as [illegible]

Ammonic Sulphate [illegible] may be obtained [illegible] phides of [illegible] ammonic sulphate [illegible]

Tests.—All the salts of ammonia [illegible] or volatilized at a high temperature [illegible] less soluble; the acid tartrate, and the double chloride of ammonium and platinum, being the least so. They therefore give, with platinic chloride and with tartaric acid, the same reactions as potassium salts. The precipitates are similar in constitution.

Ammonia may be recognised with certainty in the presence of all other bodies by its being set free, in a gaseous state, by the action of any of the caustic alkalies or alkaline earths (lime, $CaHo_2$, is the best for this purpose); by its odour, and the white fumes produced when a glass rod moistened with hydrochloric acid, HCl (not fuming acid), is brought near the compound acted on as above.

When a solution of mercuric iodide in potassic iodide (Nessler's test) is added to a solution containing an ammoniacal salt or free ammonia, to which caustic potash has been added, a brown precipitate or coloration is produced: this is by far the most delicate test for ammonia.

$$\underset{\text{Ammonia.}}{NH_3} + \underset{\text{Mercuric iodide.}}{2HgI_2} = \underset{\text{Dimercuric-ammonic iodide.}}{NHg_2I} + \underset{\text{Hydriodic acid.}}{3HI.}$$

MODES OF CHEMICAL ACTION.

In the preparation of the elementary and compound bodies which have already been described, the student will have noticed that the results were obtained by distinct chemical processes. Dr. Frankland reduces the number of these processes to five, and calls them *modes of chemical action.* Every chemical change of matter may be referred to one or other of these five modes of action :—

1. A direct combination of elements or compounds with each other.
2. A displacement of one element or compound radical in a body by another element or compound radical.
3. A mutual exchange of elements or compound radicals in two or more bodies.
4. A rearrangement of the elements or compound radicals already contained in a body.
5. The resolution of a compound into its elements, or into two or more less complex compounds.

The combustion of the metals in oxygen and chlorine are good examples of the first mode. The preparation of hydrogen, by acting on an acid with a metal, are instances of the second mode. The third mode is illustrated by the production of a salt by an acid acting on a base. There are no good illustrations in inorganic chemistry of the fourth mode ; with the exception, perhaps, of the preparation of ozone. The preparation of oxygen from mercuric oxide and potassic chlorate, and the decomposition of water and hydrochloric acid by the galvanic current, are good examples of the fifth mode of chemical action.

THE COMPOSITION OF GASES.

EXPERIMENTS AND ILLUSTRATIONS.

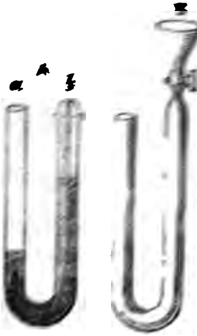


Fig. 10.

Ure's Eudiometer.—For many experiments Ure's eudiometer is a convenient instrument. It consists of a bent glass tube, closed at one end and open at the other, as shown at A, Fig. 10. The closed arm is graduated, and has two platinum wires fused into the glass, on opposite sides of the tube, in such a manner as to leave a space of about one-tenth of an inch between their points for the passage of an electric spark. The graduated arm is usually capable of containing about 40 cubic centimetres (about 2·5 cubic inches).

Composition of the Atmosphere.—This is very readily ascertained by Ure's eudiometer:—About 20 c.c. of air should be secured in the arm *b* by pouring mercury into *a*, closing the top by the thumb, and tipping the instrument so as to allow bubbles to escape from *b* into *a* till the required quantity remains. The next thing is to read off accurately the amount of air thus confined. To do so the mercury should be made to stand at the same height in each arm; this is accomplished by pouring mercury into *a*, or withdrawing it by thrusting a rod into this arm, as the case may require. Having read off the exact amount of air (say 21·5 c.c.), fill the arm *a* with mercury, and carefully closing its end with the thumb, invert it in the mercurial trough, and introduce an excess of hydrogen* (say about 10 c.c.). Again close the open end with the thumb, and pass all the gases to the closed arm; level the mercury, and read off the total volume of the mixed gases (say 31·5). Mercury should, if necessary, be poured into *a* until the distance from its surface to the top of the arm is about equal to the distance from the surface of mercury in *b* to the bend of the tube. This will prevent any gas passing the bend during the explosion. Firmly close the top *a* by the thumb, and pass an electric spark between the platinum wires. The air confined by the thumb will serve as a cushion to moderate the bursting effect of the explosion. After cooling the mercury must again be levelled, and the residuary gases read off (say 18·1).

Calculation.—Loss by explosion = 31·5 — 18·1 = 13·4. This loss is due to the formation of water by the union of hydrogen and oxygen; and as it has been explained that two volumes of hydrogen unite with one volume of oxygen, therefore one-third of this loss must have been the quantity of oxygen in the amount of air operated upon:—

$$\text{Amount of O in } 21{\cdot}5 \text{ c.c. of air} = \frac{13{\cdot}4}{3} = 4{\cdot}46 \text{ c.c.}$$

The per-centage composition is thus found:—

As 21·5 : 100 : : 4·46 : x. Therefore x = 20·74.

Per-centage of oxygen	=	20·74
„ nitrogen	=	79·26
		100·00

---

* By excess is meant more than enough for the purpose for which it is required, viz., to unite with the oxygen. 21·5 c.c. of air contain about 4·3 c.c. of oxygen, which will unite with 8·6 c.c. of hydrogen; 10 c.c. is therefore an excess.

Composition of Water.—This has been proved analytically by electrical decomposition, but it may also be proved *synthetically* by the eudiometer.

First introduce about 8 c.c. of oxygen and read off accurately; then add more than double this amount of hydrogen and ascertain the total quantity of mixed gases. Explode, cool, and read off the residue, which proves to be pure hydrogen.

Example.—

Amount of oxygen used . = 8·5 c.c.
„ mixed gases . = 28·4 „
Residual H after explosion = 2·9 „
Loss by explosion . . = 25·5 „
But the loss due to oxygen . . = 8·5
Therefore loss due to H = 25·5 − 8·5 = 17

Hence 17 c.c. of H unite with 8·5 c.c. of O, or two volumes of hydrogen unite with one volume of oxygen to form water.

To prove that these three volumes form only two of water vapour, the whole apparatus may be placed in an oil bath having a temperature above 100° C. After corrections for temperature, which will be explained further on, the amount of this vapour will be found to be 17 c.c., or two volumes.

this assists its introduction, and also prevents the escape of the confined gas.

The sodium of the amalgam readily acts on the hydrochloric acid gas, liberating pure hydrogen. Only a portion of the amalgam must be allowed to enter, or the air will rush in and spoil the experiment. When the action is completed it will be found that one half the original volume of gas remains, and that it is pure hydrogen: *therefore two volumes of* HCl *contain one volume of* H. It only remains to be proved that these two volumes contain one volume of chlorine. The gases cannot be mixed and exploded, because chlorine attacks mercury and is soluble in water.

Take a solution of HCl and thoroughly saturate it with chlorine gas, then subject the liquid to the action of an electric current. Allow the action to proceed some time, then collect the gases in separate tubes. It will be found that H and Cl come off in equal volumes.

Therefore in the original experiment the two volumes of hydrochloric acid gas contained one volume of hydrogen and one volume of chlorine.

Hydrobromic and hydriodic acids may also be experimented on in the eudiometer, as described above.

Composition of Ammonia Gas.—The constituent gases of ammonia refuse to unite directly, and neither of them can be conveniently withdrawn, as in the case of hydrochloric acid gas. A different method is therefore adopted.

Take a graduated tube, about 18 inches or 2 feet long, fill it with chlorine gas, and invert it in a deep vessel containing solution of ammonia. On shaking the tube up and down rapid absorption takes place, till a gas remains, occupying one-third the volume of the chlorine. Examination proves this gas to be pure nitrogen. (See the action of Cl on solution of ammonia, page 94.) In this experiment three volumes of chlorine were required to liberate one volume of nitrogen, and we know that each volume of this chlorine united with *one* volume of hydrogen; therefore the *three* volumes of chlorine united with *three* volumes of hydrogen, and by so doing liberated *one* volume of nitrogen from its combination with the hydrogen: hence one volume of nitrogen unites with three volumes of hydrogen to form ammonia.

We have already seen that two volumes of ammonia gas break up into one volume of nitrogen and three of hydrogen, under the influence of powerful electric sparks.

INFLUENCE OF TEMPERATURE AND PRESSURE ON THE VOLUMES OF GASES.

Temperature.—Nearly all bodies expand or increase in bulk when heated. Gases and vapours expand much more than either liquids or solids. The amount of expansion has been carefully ascertained, and it is found to be the same for all gases and vapours at efficient of expansion, and the co-efficient is the same under all pressures.

As the co-efficient is calculated from the volume of the gas at zero (0° C.), this is called the *normal* or standard temperature.

The following examples will illustrate the use of the co-efficient in calculations of gaseous volumes:—

273 litres of gas at 0° C., if raised to—

1° C. will measure 274 litres (= 273 + $\frac{1}{273}$ of 273).
2° C. „ 275 „ (= 273 + $\frac{2}{273}$ of 273).
3° C. „ 276 „ (= 273 + $\frac{3}{273}$ of 273).
10 C. „ 283 „ (= 273 + $\frac{10}{273}$ of 273).

And if reduced to—

− 1° C. will measure 272 litres (= 273 − $\frac{1}{273}$ of 273).
− 2 C. „ 271 „ (= 273 − $\frac{2}{273}$ of 273).
− 20° C. „ 253 „ (= 273 − $\frac{20}{273}$ of 273).

It will readily be seen that the above data may be used in calculations of changes of volume.

Example 1.—100 litres of gas are measured off at 0° C.: what will they measure when the temperature is raised to 50° C.?

The volume will evidently bear the same relation to 100 that 273 + $\frac{50}{273}$ of 273 does to 273.

$$\begin{array}{ccccccc} \text{At } 0^\circ. & & \text{At } 50^\circ. & & \text{Litres at } 0^\circ. & & \text{Litres at } 50^\circ. \\ \therefore \text{ as } 273 & : & 323 & :: & 100 & : & x \end{array}$$

$$x = \frac{323 \times 100}{273} = 118{\cdot}31 \quad \text{No. of litres at } 50^\circ \text{ C.}$$

Example 2.—100 litres of gas are measured off at 50° C.: what will they measure if reduced to 20° C.?

$$\begin{array}{ccccccc} \text{Vol. at } 50^\circ. & & \text{Vol. at } 20^\circ. & & \text{Litres at } 50^\circ. & & \text{Litres at } 20^\circ. \\ \text{As } 273 + 50 & : & 273 + 20 & :: & 100 & : & x \end{array}$$

$$x = \frac{293 \times 100}{323} = 90{\cdot}71 \text{ No. of litres at } 20^\circ \text{ C.}$$

Example 3.—50 litres of gas are measured off at 30° C. : what will they measure if the temperature be reduced to – 20° C.?

$$\begin{array}{lcccc} & \text{Vol. at } 30^\circ. & \text{Vol. at} - 20^\circ. & \text{Litres at } 30^\circ. & \text{Litres at} - 20^\circ. \\ \text{As} & 273 + 30 : & 273 - 20 & :: \quad 50 \quad : & x \end{array}$$

$$x = \frac{253 \times 50}{303} = 41{\cdot}74 \text{ No. of litres at} - 20^\circ.$$

The degrees of temperature employed in this book are those of the Centigrade thermometer, but in England temperature is very frequently given in degrees Fahrenheit, and in some Continental countries in those of Reaumur; it is therefore necessary that the student should be able to convert the degrees of one scale into those of another: for this purpose the following equations will assist him:—

$$C. = \frac{5(F. - 32).}{9}$$

Example 5.—Convert 14° F. into degrees Centigrade.

$$C. = \frac{5(14-32)}{9} = \frac{-90}{9} = -10° C.$$

The student should not neglect to take the *minus* sign into account in examples like No. 5.

Pressure.—The volume of a gas alters with every change of external pressure which it experiences; the amount of such change is a constant quantity, and is always calculated by the *law of Boyle and Mariotte*, which states that "the volume of a gas is inversely as the pressure which it sustains, if the temperature remain the same."

The *normal* or standard pressure is that of a column of mercury measuring 760 millimetres. The mercurial barometer is the instrument used to determine this and all other pressures in calculations of gaseous volumes.

The following examples will serve to illustrate the mode of applying Boyle and Mariotte's law for the calculations of gaseous volumes:—

Example 6.—100 litres of gas are measured off at the normal pressure: what will they measure if the pressure be reduced to 750 m.m., the temperature remaining the same?

The volume will be inversely as the pressure.

	M.m.		M.m.		Lit. at 760 m.m.		Lit. at 750 m.m.
∴ As	750	:	760	::	100	:	x

$$x = \frac{760 \times 100}{750} = 101 \cdot 3 \text{ litres at } 750 \text{ m.m.}$$

When there is a change both in the temperature and pressure, the resulting volume may be arrived at either by calculating for the temperature first and fo the pressure afterwards, or both together.

Example 7.—100 litres of gas are measured off a 30° C. and 765 m.m. pressure: what will they measure at the normal temperature and pressure?

	Vols. at 30°.		Vols. at 0°.		Litres. at 30°.		Litres at 0° and 765 m.m. pressure.
Correction for temperature	273 + 30	:	273	::	100	:	x

			Litres at 0° and 765 m.m.		Litres at 0° and 762 m.m. pressure.
Correction for pressure	760 : 765	::	x	:	x'

$$\therefore x' = \frac{273 \times 100}{303} \times \frac{765}{760} = 90{\cdot}69$$ No. of litres at 0° C. and 760 m.m. pressure.

The Crith.—This is the name which Dr. Hofmann has given to the weight of a litre of hydrogen at the normal temperature and pressure. It [redacted] standard unit of volume lations

Example 8.—What is the weight of 50 litres of nitrogen at the normal temperature and pressure?

Grammes.
1 litre of nitrogen weighs 14 criths = 14 × ·0896

Grammes.
∴ 50 litres do. 50 × 14 criths = 50 × 14 × ·0896 = 6·272 grammes.

Example 9.—How many litres will 50 grammes of oxygen measure at the normal temperature and pressure?

Grammes of oxygen.		Grammes of oxygen.		Litre.		Litres.
As 16 × ·0896	:	50	::	1	:	x

$$x = \frac{50}{16 \times \cdot 0896} = 34\cdot 87 \text{ No. of litres of oxygen.}$$

The densities of all compound gases are equal to half their molecular weights. This arises from the fact that all molecules at the same temperature and pressure occupy the same volume, and that this volume is equal to two volumes of hydrogen. One volume of hydrogen is the unit for comparison of densities, and therefore one volume only of the compound must be weighed for its density, but this one volume is half the molecule.

$$\text{Density of HCl} = \frac{1 + 35\cdot 5}{2} = 18\cdot 25$$

$$\text{Do.} \quad OH_2 \text{ (vapour)} = \frac{16 + 2}{2} = 9$$

$$\text{Do.} \quad NH_3 = \frac{14 + 3}{2} = 8\cdot 5$$

And converting these densities into criths, we have—

1 litre of	HCl gas . .	18·25 criths.
1 „	OH_2 (vapour) . .	9 „
1 „	NH_3 . . .	8·5 „

The weight, therefore, of a litre of any compound gas or vapour is equal to half the molecular weight in criths.

Example 10.—What is the weight of 10 litres of hydrobromic acid gas at the normal temperature and pressure?

$$\text{Weight of 1 litre of HBr} = \frac{1 + 80}{2} = 40{\cdot}5 \quad \text{criths.}$$

$$\text{,, } \quad 10 \text{ litres ,, } \quad = 40{\cdot}5 \times 10 \text{ ,,}$$

And 1 crith = ·0896 gramme.

Grammes.

$$\therefore 40{\cdot}5 \times 10 \text{ criths} = {\cdot}0896 \times 40{\cdot}5 \times 10 = 36{\cdot}288.$$

Example 11.—How many litres of HCl gas at the normal temperature and pressure can be obtained by acting on 100 grammes of sodic chloride with sulphuric acid?

The reaction is shown by the following equation:—

$$2NaCl + SO_2Ho_2 = 2HCl + SO_2Nao_2.$$

$$\begin{array}{lccc} \text{Grammes HCl.} & \text{Grammes HCl.} & \text{Litre.} & \text{Litres.} \\ \therefore \text{As } 18\ 25 \times {\cdot}0896 & : \ 62{\cdot}39 \ :: & 1 & : \ x \end{array}$$

$$x = \frac{62{\cdot}39}{18{\cdot}25 \times {\cdot}0896} = 38{\cdot}15 \text{ litres of hydrochloric acid gas.}$$

Dr. Williamson has adopted a different standard volume, which may be used with advantage in many calculations. It is the volume occupied by 1 gramme of hydrogen at 0° C. and 760 m.m., which equals 11·19 litres. The same space is occupied by 35·5 grammes of chlorine, 16 grammes of oxygen, 14 grammes of nitrogen, 8 5 grammes of ammonia gas, and 18·25 of hydrochloric acid.

By this standard the 62·39 grammes of HCl can be converted into litres as follows:—

$$\begin{array}{lccc} \text{Gram. HCl.} & \text{Gram. HCl.} & \text{Litres.} & \text{Litres.} \\ \text{As } 18{\cdot}25 & : \ 62{\cdot}39 \ :: & 11{\cdot}19 & : \ x \end{array}$$

$$x = \frac{62{\cdot}39 \times 11{\cdot}19}{18{\cdot}25} = 38{\cdot}25 \text{ No. of litres of HCl.}$$

The two results are not exactly the same, because 11·19 litres is not the exact measurement of 1 gramme of hydrogen, and ·0896 is not the exact weight of a litre of hydrogen, but the results approximate sufficiently for all practical purposes.

HEXAD ELEMENTS.

Sulphur—S_2.

Atomic weight and sp. gr. of vapour, 32. Molecular weight, 64. Molecular volume, normal at 1000°, but only one-third of the normal molecular volume at temperatures near its boiling point. 1 litre of its vapour weighs 32 criths. Atomicity, VI., IV., and II.

Occurrence in Nature.—Much of the sulphur employed in this country comes from Sicily and other volcanic districts, where it is met with in the free or uncombined state. In combination with iron, antimony, zinc, lead, copper, and other metals, as sulphides, it occurs abundantly in nature; in union with oxygen it is found combined with calcium, barium, strontium, &c., as sulphates; it is present in small quantities in all animal and vegetable substances. It exists in the albumen of eggs. The blackening of silver spoons used with eggs is owing to the formation of the sulphide of silver.

Properties.—At ordinary temperatures this element is a pale yellow crystalline solid.

Sulphur melts at 110° C., becoming a thin amber-coloured liquid; when heated to 220° it thickens like treacle. In this state it may be poured from a flask into cold water, when it retains for some time a soft, flexible condition, like gutta-percha; this is the amorphous condition of sulphur. If it be raised to a temperature of 420° it boils; and if the vapour be conducted into a cold chamber it condenses in crystals known as *sublimed sulphur*, or *flowers of sulphur*. The *roll sulphur* or *brimstone* of commerce is formed by running it when in a melted state into wooden moulds.

Sulphur is readily inflammable, taking fire when heated to about 240°; it burns with a blue flame, In these sulphides the sulphur is almost always a diad.

Sulphur assumes several allotropic forms, of which the following are the principal :—

1. Octahedral crystals, as in native sulphur, and by evaporating solution of sulphur in carbonic disulphide. This form is soluble in carbonic disulphide.

2. Prismatic crystals, as in roll sulphur.

3. Plastic or amorphous, prepared as stated above ; it is insoluble in carbonic disulphide.

4. Powder of a yellow or orange colour, also insoluble in carbonic disulphide.

Sulphur is used extensively in the manufacture of gunpowder, matches, and sulphuric and sulphurous acids ; also in medicine.

Compounds of Sulphur with Basylous Elements.

Sulphuretted Hydrogen, or Hydrosulphuric Acid, SH_2, is a gas which is evolved from volcanoes, and is found in solution in spa waters, such as those at Harrogate, and in waters which contain both organic matter and sulphates. It is a constant product of the decay of animal and vegetable matter.

Preparation.—1. It is usually prepared by acting on ferrous sulphide with dilute sulphuric acid :—

$$FeS + SO_2Ho_2 = SH_2 + SO_2Feo''.$$

The gas is liberated without heat, and an ordinary gas bottle (*Fig.* 11) may be used. It should be collected over warm water. A solution of the gas may be prepared by passing it into cold water (*Fig.* 12), contained in a bottle, which should be only half full. The bottle should be closed with a stopper from

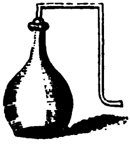


FIG. 11.

time to time, and the contents well shaken up. In this manner the water becomes more quickly saturated with the gas.

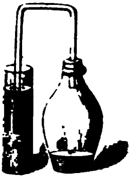


FIG. 12.

Hydrochloric acid may be substituted for the sulphuric in this process, but it does not answer so well.

The ferrous sulphide is made by holding a roll of sulphur in contact with a piece of wrought iron heated to a white heat; the fluid sulphide should be allowed to drop into a bowl of water. It may also be prepared by projecting a mixture of iron borings and powdered sulphur into a red-hot fireclay crucible, in small portions at a time.

2. Antimonious sulphide is heated with strong hydrochloric acid in a gas bottle (*Fig.* 11):—

$$Sb_2S_3 + 6HCl = 2SbCl_3 + 3SH_2.$$

The gas obtained by this process is pure.

3. By heating sulphur in hydrogen gas.

Properties.—Sulphuretted hydrogen, at the common temperature, is a transparent colourless gas, Water, at the ordinary temperature, dissolves from two to three times its own volume of this gas; the solution is used as a test. It soon spoils by absorption of oxygen, and is best prepared as wanted. The gas has the properties of an acid, reddens litmus, and is called *hydrosulphuric acid*. A piece of white blotting-paper dipped in a solution of plumbic acetate (sugar of lead) is immediately blackened by the smallest trace of this gas. This method is frequently adopted to test the purity of coal gas. Silversmiths usually burn their gas outside, owing to the sulphuretted hydrogen, usually present in small quantities in coal gas, blackening the silver by the union of the sulphur with that metal. Its specific gravity is $\frac{2 + 32}{2} = 17$. One litre of it therefore weighs 17 criths. Its molecular volume is normal.

Uses.—Sulphuretted hydrogen is invaluable to the analyst as a general or group reagent. It precipitates the following metals as sulphides from acid solutions of any of their salts:—Pb, Cu, Bi, Cd, Hg, Sn, Sb, Au, and Pt; and the following metals from solutions of their salts previously made alkaline by addition of ammonic hydrate:—Fe, Co, Ni, Mn, and Zn as sulphides, Cr and Al as hydrates. Some of the metals can be recognised at once by the colour of the precipitated sulphide.

The following equations will illustrate the action of the reagent:—

$$\underset{\substack{\text{Cupric}\\\text{sulphate.}}}{SO_2Cuo''} + \underset{\substack{\text{Sulphuretted}\\\text{hydrogen.}}}{SH_2} = \underset{\substack{\text{Sulphuric}\\\text{acid.}}}{SO_2Ho_2} + \underset{\substack{\text{Cupric}\\\text{sulphide.}}}{CuS}.$$

$$\underset{\substack{\text{Bismuthous}\\\text{chloride.}}}{2BiCl_3} + 3SH_2 = 6HCl + \underset{\substack{\text{Bismuthous}\\\text{sulphide.}}}{Bi_2S_3}.$$

$$\underset{\text{Ferrous sulphate.}}{\mathbf{S}O_2\mathbf{Fe}o''} + \underset{\text{Ammonic hydrate.}}{2AmHo} + \mathbf{S}H_2 = \underset{\text{Ammonic sulphate.}}{\mathbf{S}O_2Amo_2} + \underset{\text{Ferrous sulphide.}}{\mathbf{FeS}} + 2\mathbf{O}H_2.$$

Sulphuretted hydrogen acts as a reducing agent in some cases; the following is a familiar example:—

$$2\underset{\text{Ferric chloride.}}{\mathbf{Fe}'''_2Cl_6} + 2\mathbf{S}H_2 = 4\underset{\text{Ferrous chloride.}}{\mathbf{Fe}''Cl_2} + 4HCl + \mathbf{S}_2.$$

Hydrosulphyl, $\mathbf{S}_2H_2$, or Hs_2, (persulphide of hydrogen).—This substance is prepared by mixing a solution of calcic disulphide with HCl:—

$$\mathbf{CaS}_2 + 2HCl = \mathbf{S}_2H_2 + \mathbf{Ca}Cl_2.$$

It is the analogue of hydroxyl, and performs the same functions in sulphydrates that hydroxyl does in hydrates, *e. g.*:—

Zincic hydrate, $\mathbf{Zn}Ho_2$; zincic sulphydrate, $\mathbf{Zn}Hs_2$.

Carbonic Disulphide, $\mathbf{C}\mathbf{S}''_2$.—This substance is also known as *bisulphide of carbon*. It is a heavy volatile liquid, boiling at 45°, and having a sp. gr. of 1·26. The molecular volume of its vapour is normal.

by passing the vapour of sulphur through an iron cylinder containing charcoal heated to redness. The gaseous product is passed into a well-cooled condenser, where it liquefies:—

$$C + S = CS_2.$$

Compounds of Sulphur with Oxygen and Hydroxyl.

These compounds are somewhat numerous, and some of them are of great importance in the arts.

Sulphurous Anhydride, $S''O_2$.—This compound may be prepared in a variety of ways.

1. By burning sulphur in air or oxygen:—

$$S_2 + 2O_2 = 2SO_2.$$

2. By reducing sulphuric acid with copper, mercury, or charcoal:—

$$2SO_2Ho_2 + Cu = SO_2Cuo'' + 2OH_2 + SO_2.$$
$$2SO_2Ho_2 + C = 2OH_2 + 2SO_2 + CO_2.$$

The real reducing agent in the first of these examples is the nascent hydrogen displaced by the copper. Process 2 is conducted in a gas bottle (*Fig.* 11). The gas must be collected over mercury or by displacement, as it is soluble in water.

Properties.—A transparent, colourless gas, having the suffocating odour of burning sulphur. It liquefies at —10°, or under a pressure of two atmospheres at ordinary temperatures. The rapid evaporation of the liquid produces great cold. Water at 0° absorbs and combines with 69 times its volume of this gas, producing a solution of sulphurous acid:—

$$OH_2 + SO_2 = \underset{\text{Sulphurous Acid.}}{SOHo_2.}$$

The gas passed into water in which metalli hydrates or carbonates are dissolved or suspende converts them into sulphites; *e.g.*,—

$$\text{ONaH} + \text{SO}_2 = \underset{\substack{\text{Hydric-sodic}\\ \text{sulphite.}}}{\text{SOHoNao.}}$$

$$\text{COCao}'' + \text{SO}_2 = \underset{\substack{\text{Calcic}\\ \text{sulphite.}}}{\text{SOCao}''} + \text{CO}_2.$$

This gas has bleaching properties, and is used for bleaching straw used in plaiting, and woollen goods. The plaiters in Bedfordshire suspend their plait in a tea-chest, and set fire to some sulphur in an iron dish at the bottom; the top of the chest is then covered over with a sack. If violets, or any coloured flowers, be suspended in a jar of this gas, the colour is entirely removed. Litmus paper dipped in a solution of sulphurous acid is first reddened, then bleached. It has also the power of arresting animal decomposition, and is therefore useful as a disinfectant and antiseptic.

Sulphurous Acid, $SOHo_2$.—We have seen this substance is produced when combines with [redacted]

Sulphuric Anhydride, SO_3, is a solid, and may be obtained by gently heating Nordhausen sulphuric acid in a retort, provided with a receiver, surrounded by a freezing mixture. The vapour, on reaching the receiver, is condensed into white silky crystals.

It may also be prepared by heating sulphuric acid with phosphoric anhydride in a retort. The sulphuric acid is deprived of water by the phosphoric anhydride, thus :—

$$SO_2Ho_2 + \underset{\substack{\text{Phosphoric}\\ \text{anhydride.}}}{P^{v}_2O_5} = \underset{\substack{\text{Metaphosphoric}\\ \text{acid.}}}{2PO_2Ho} + \underset{\substack{\text{Sulphuric}\\ \text{anhydride.}}}{SO_3}.$$

Sulphuric anhydride has no acid properties; but when thrown into water it hisses like red-hot iron, combining with the water, and forming sulphuric acid. On exposure to the air it liquefies by the absorption of moisture.

Sulphuric Acid, SO_2Ho_2.—This is a most important compound. It is extensively used in the manufacture of soda from common salt, for bleaching, for dissolving indigo, dyeing, calico-printing, and numerous other purposes. It has been known from the fourteenth century.

Manufacture.—In Great Britain this manufacture is of great national importance, and is carried on to a vast extent.

The method adopted in the manufacture of sulphuric acid depends on the formation of sulphurous anhydride, and its subsequent conversion into sulphuric acid by the combined action of oxygen and moisture. It is therefore entirely a synthetical process.

Sulphur, either in the form of native sulphur or *iron pyrites*, is burnt in a furnace freely supplied with air, whereby sulphurous anhydride is formed (*Fig.* 13),—

$$S_2 + 2O_2 = 2SO_2,$$

which is passed into a chamber made of sheet lead, and supported by a framework of wood. An iron pot or crucible hangs over the fire, containing a mixture of sodic nitrate and sulphuric acid, from which are evolved mixed fumes of nitric acid and nitric peroxide.[*] These mingle with the sulphurous anhydride, and pass with it into the chamber, through a short, wide pipe. Jets of steam also play into the room, and the floor is covered with water to the depth of a few inches. Four substances are therefore essential for the production of sulphuric acid, viz., sulphurous acid, air, water, and nitric acid.

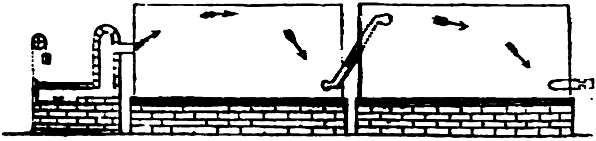


FIG. 13.

Sulphurous acid has already been stated to be a reducing agent, becoming sulphuric acid at the expense whilst at the same time the sulphurous anhydride becomes sulphuric acid, thus :—

$$\underset{\substack{\text{Nitric}\\ \text{acid.}}}{2NO_2Ho} + \underset{\substack{\text{Sulphurous}\\ \text{anhydride.}}}{SO_2} = \underset{\substack{\text{Sulphuric}\\ \text{acid.}}}{SO_2Ho_2} + \underset{\substack{\text{Nitric}\\ \text{peroxide.}}}{N_2O_4}.$$

The nitric peroxide, however, does not remain in this state, but suffers further deoxidation until it becomes reduced to nitric oxide, thus :—

$$\underset{\substack{\text{Nitric}\\ \text{peroxide.}}}{N_2O_4} + 2SO_2 + 2OH_2 = 2SO_2Ho_2 + \underset{\substack{\text{Nitric}\\ \text{oxide.}}}{N_2O_2}.$$

The nitric oxide suffers no further reduction, but in contact with the oxygen of the air it becomes once more nitric peroxide (see Properties of Nitric Oxide),—

$$N_2O_2 + O_2 = N_2O_4;$$

which is again destined to be reduced by the action of a fresh supply of sulphurous anhydride. Thus the nitrogen compounds derived from the nitric acid fumes are capable of acting again and again as carriers of oxygen from air to the sulphurous acid, and a limited quantity would act continuously and to an indefinite extent, were it not that a portion is continually escaping along with the atmospheric nitrogen into the tall chimney with which the chambers are connected.

The above is the simplest view that can be taken of the reactions which occur in the vitriol chamber. Many chemists believe that they are much more complex. Thus, probably, nitrous anhydride is one of the nitrogen products at work,—

$$4NO_2Ho + 4SO_2 + 2OH_2 = 4SO_2Ho_2 + 2N_2O_3;$$

Nitric acid. Sulphurous anhydride. Sulphuric acid. Nitrous anhydride.

and the nitrous anhydride thus formed is either converted by water into a mixture of nitric acid and nitric oxide (both of which we have seen are capable of oxidizing sulphurous acid to sulphuric acid), or it acts directly as the oxidizer of sulphurous anhydride. The following are equational expressions of both reactions :—

$$3N_2O_3 + OH_2 = 2NO_2Ho + 2N_2O_2.$$

Nitrous anhydride. Water. Nitric acid. Nitric oxide.

$$N_2O_3 + SO_2 + OH_2 = SO_2Ho_2 + N_2O_2.$$

Nitrous anhydride. Sulphurous anhydride. Sulphuric acid. Nitric oxide.

A peculiar white crystalline compound, to which the following formula has been assigned, is observed to be produced in the vitriol chamber :—

$$\left\{\begin{matrix} SO_2(NO_2)' \\ O \end{matrix}\right.$$ White crystalline compound of Brü-

of vitriol;" or it is further concentrated in platinum or glass vessels up to a density when cold of 1·845 or 1·85, and is then known in commerce as "double oil of vitriol." This is the strongest English acid.

Sulphuric acid is also formed when water is added to sulphuric anhydride, or hydroxyl to sulphurous anhydride, and it is slowly produced when sulphurous anhydride is exposed to the action of the air.

Properties.—Sulphuric acid is a colourless oily liquid of intensely acid reaction. It boils at 327° C., and distils unchanged. Organic matter is rapidly charred and destroyed by it. When mixed with water great heat is evolved; it has a great affinity for water. Gases are frequently dried by being transmitted through this acid.

When the vapour of sulphuric acid is passed over red-hot platinum it breaks up into SO_2, OH_2, and oxygen. With the aid of heat it decomposes all salts containing acids more volatile than itself; it is therefore extensively used in the preparation of other acids (see Preparation of Nitric and Hydrochloric Acids). It is a dibasic acid, and forms *sulphates* with bases, most of which are soluble in water, but none of them in alcohol.

The molecular volume of sulphuric acid vapour is abnormal, being 4 volumes, owing probably to dissociation.

Test.—Baric chloride or nitrate in dilute solution gives a white precipitate of baric sulphate with dilute sulphuric acid, or a solution of a sulphate; the precipitate is insoluble in nitric acid.

Nordhausen Sulphuric Acid, $\left\{\begin{array}{l} SO_2Ho \\ O \\ SO_2Ho \end{array}\right.$, is a kind of acid manufactured on the Continent from crystallized ferrous sulphate, $SOHoFeo'',6OH_2$, which is first well dried in order to drive off most of the water of crystallization, and then distilled at a red heat in an earthenware retort.

It fumes in moist air, has a sp. gr. of 1·9, and is used for dissolving indigo, and for the preparation of sulphuric anhydride.

Hyposulphurous Acid, $SS''OHo_2$, is not known in the free state. In combination with metals it forms hyposulphites, for the preparation of which the student is referred to sodic hyposulphite, the most important of these salts. The acid may be regarded as sulphuric acid in which one of the atoms of oxygen is replaced by sulphur, and it is obtained by the sulphuration of sulphurous acid, and is sometimes called sulphosulphuric acid.

Four other oxy-acids of sulphur are known, but as they are not at present of great practical importance, their names and formulæ only will be given. As the usual nomenclature has already been exhausted in the names of the foregoing acids, the Greek noun for sulphur (θειον) is made use of for the names of these

HALOGENS.

Chlorine, Bromine, Iodine, Fluorine.

Chlorine and its compounds have already been described. The members of this group possess a remarkable resemblance to one another, and there is a curious gradation in their atomic weights, specific gravities, bleaching properties, and physical condition. They combine directly with metals to form salts, and with hydrogen they form hydracids.

By a careful study of the halogens a great similarity will be observed in their chemical characters, the chief differences being in the *intensity* of their chemical action. They all unite with hydrogen, forming acids, and produce similar salts by the substitution of metals for this hydrogen. The similarity between the first three is very remarkable. They all unite with silver, forming curdy precipitates, insoluble in nitric acid.

MONAD ELEMENTS (*continued*).

Bromine—Br_2.

Atomic wt. and sp. gr. of vapour, 80. Molecular weight, 160. Molecular volume, normal. 1 litre of vapour weighs 80 criths. Atomicity, I.

Occurrence in Nature.—As magnesic bromide, $MgBr_2$, in sea water. Abundant in the waters of the Dead Sea.

Preparation.—It is obtained from *bittern*, the liquor which remains after most of the salts have been removed from sea water by evaporation. A current of chlorine gas is passed through the bittern which liberates the bromine, which is combined chiefly with magnesium, and the bromine is then taken up by ether:—

$$MgBr_2 + Cl_2 = MgCl_2 + Br_2.$$

Bromine may also be made by acting on a bromide with manganic oxide and sulphuric acid. The reaction is analogous to that which liberates chlorine from a chloride under like conditions.

Properties.—A deep red heavy liquid, giving off offensive red fumes: the name bromine is derived from βρῶμος (a stench). The liquid boils at 63° C., and freezes at—20° C. It dissolves in water slightly, but abundantly in ether and carbonic disulphide. Bromine has bleaching properties when associated with water, but in a less degree than chlorine. Powdered antimony and arsenic burn in bromine vapour, and most metals are attacked by it and converted into bromides. It is much used in the production of organic compounds; it is also used largely in photography.

Hydrobromic Acid, HBr.—This acid may be prepared by heating a bromide with phosphoric acid in a flask with delivery tube; the liberated gas may then be conducted into water, in which it freely dissolves:—

$$3KBr + POHo_3 = POKo_3 + 3HBr.$$

Potassic	Phosphoric	Potassic	Hydrobromic

$$\underset{\substack{\text{Potassic}\\ \text{hydrate.}}}{6KHo} + \underset{\text{Bromine.}}{3Br_2} = \underset{\substack{\text{Potassic}\\ \text{bromide.}}}{5KBr} + \underset{\substack{\text{Potassic}\\ \text{bromate.}}}{\left\{\begin{matrix}OBr\\ OKo\end{matrix}\right.} + 3OH_2.$$

The bromate is separated from the bromide by crystallization. The potassic bromate is converted into baric bromate by adding solution of baric chloride to it, and separating the two products by crystallization, and the baric bromate is then treated with sulphuric acid, which liberates bromic acid, and insoluble baric sulphate precipitates, which may be filtered off:—

$$\underset{\substack{\text{Baric}\\ \text{bromate.}}}{\left\{\begin{matrix}OBr\\ O\\ Bao''\\ O\\ OBr\end{matrix}\right.} + \underset{\substack{\text{Sulphuric}\\ \text{acid.}}}{SO_2Ho_2} = 2\underset{\substack{\text{Bromic}\\ \text{acid}}}{\left\{\begin{matrix}OBr\\ OHo\end{matrix}\right.} + \underset{\substack{\text{Baric}\\ \text{sulphate.}}}{SO_2Bao''}.$$

Bromic acid closely resembles chloric acid. It is easily decomposed. Alkaline bromates, when heated, yield oxygen and bromides.

Iodine—I_2.

Atomic weight and sp. gr. of vapour, 127. Molecular weight, 254. Molecular volume, normal. 1 litre of vapour weighs 127 criths. Atomicity, I.

Occurrence in Nature.—It is found principally in sea plants, which derive it from sea water. It also occurs in mineral waters.

Preparation.—The fused ashes of seaweeds, called *kelp*, are treated with boiling water, which dissolves out considerable quantities of sodic carbonate and chloride and other salts, and potassic iodide; the sodic salts, &c., are crystallized out as far as possible, and the potassic iodide remains in the mother liquor, which is then treated with manganic oxide and sulphuric acid, and distilled; the iodine distils over, the reaction being analogous to that which yields chlorine and bromine by similar treatment. This process may be imitated by placing in a retort connected with a receiver a strong solution of potassic iodide, a little powdered binoxide of manganese, and sulphuric acid: on the application of heat a beautiful violet vapour passes over, which condenses in the cold part of the retort and receiver in small shining scales.

Properties.—At ordinary temperatures iodine is a solid, having a bluish black metallic appearance, like black-lead. It fuses at 107°, and boils at 180°, passing into a vapour of a splendid violet colour—the name iodine being derived from ἰώδης (violet-coloured). The sp. gr. of the solid is 4·95. It is slightly soluble in water, but freely in alcohol and ether, the solutions of which are brown. Carbonic disulphide and chloroform also dissolve it, the solutions being violet like the vapour.

Hydriodic Acid, HI.—This acid is a gas, but it may be obtained in solution by passing sulphuretted hydrogen through water in which powdered iodine is suspended. In this case the sulphur of the sulphuretted hydrogen is liberated, and the hydrogen combines with the iodine, forming hydriodic acid, $2SH_2 + 2I_2 = 4HI + S_2$. The liquid is warmed, filtered, and a colourless solution of the acid is obtained, which soon becomes of a brown colour on exposure to air, by its decomposition into free iodine and water, $4HI + O_2 = 2OH_2 + 2I_2$.

The gas is best prepared by cautiously adding a small quantity of water to phosphorous triiodide and collecting by displacement:—

$$PI_3 + 3OH_2 = POHHo_2 + 3HI.$$

The phosphorous acid remains in solution.

It may also be obtained by processes analogous to those employed in the preparation of HBr and HCl.

This gas is very heavy, its molecular weight being 128, and its density (= half molecular weight) is 64. It is decomposed by chlorine, bromine, nitrous, nitric, and sulphuric acid, and slowly by air. It forms iodides with bases, and dissolves many metals with evolution of hydrogen:—

$$Zn + 2HI = ZnI_2 + H_2.$$

Iodic Acid, $\left\{\begin{matrix}OI\\OHo\end{matrix}\right.$.—This acid is obtained from baric iodate by processes analogous to those which yield bromic acid, to which the student is referred. It may also be produced by oxidizing iodine with strong boiling nitric acid, or with chlorine water. The following equation explains the latter process:—

$$I_2 + 6OH_2 + 5Cl = 2\left\{\begin{matrix}OI \\ OHo\end{matrix}\right. + 10HCl.$$

Solution of iodic acid is readily reduced by sulphurous acid, the solution becoming brown from separated iodine.

When iodic acid is poured into a solution of hydriodic acid, iodine is liberated:—

$$\underset{\substack{\text{Iodic}\\\text{acid.}}}{\left\{\begin{matrix}OI \\ OHo\end{matrix}\right.} + \underset{\substack{\text{Hydriodic}\\\text{acid.}}}{5HI} = \underset{\text{Water.}}{3OH_2} + \underset{\text{Iodine.}}{3I_2}.$$

Iodates are formed by similar means to those which yield bromates and chlorates, which they closely resemble. Iodates, when heated, break up into iodides and oxygen, or into oxides, iodine, and oxygen.

Periodic Acid, $\left\{\begin{matrix}OI \\ O \\ OHo\end{matrix}\right.$, is produced by decomposing plumbic periodate with sulphuric acid, and decanting the solution from the insoluble plumbic sulphate. It crystallizes from its solution in deliquescent silicon and hydrogen, it has never been with certainty isolated. Its most frequent combination is with calcium, as calcic fluoride, CaF_2, or *fluor spar*, and as the mineral *cryolite*, $6NaF', Al'''_2F_6$. It exists also in a state of combination in the bones and teeth of animals. So far as it has been examined, it is said to be a gas, like chlorine.

There is no known oxygen compound of fluorine; but in combination with hydrogen it forms an acid known as *hydrofluoric acid*, HF, which, by the displacement of its hydrogen, furnishes fluorides. The compounds of fluorine are closely analogous to those of chlorine, bromine, and iodine.

Hydrofluoric Acid, HF.—To obtain this compound, the powdered calcic fluoride is heated with two or three times its weight of concentrated sulphuric acid in a platinum or lead retort connected with a receiver of the same metal, kept cool by being surrounded with a freezing mixture. A volatile colourless liquid is obtained, which emits suffocating white fumes in the air. It unites with water, and even very dilute solutions attack glass. It has, therefore, to be kept in leaden or gutta-percha bottles.

The decomposition is easily explained:—

$$\underset{\substack{\text{Calcic}\\\text{fluoride.}}}{CaF_2} + SO_2HO_2 = 2HF + \underset{\substack{\text{Calcic}\\\text{sulphate.}}}{SO_2Cao''}.$$

If the concentrated acid be dropped upon the skin, it occasions ulcers of a very malignant and dangerous kind. It is used to etch or engrave letters on glass. The glass to be operated upon is first covered with beeswax, and the lines to be etched scratched through with a needle or any sharp-pointed instrument. The glass is then exposed to the action of the vapour of hydrofluoric acid, or the acid is poured upon it in a liquid state. Where the glass is unprotected with wax it is corroded. It is in this manner that thermometers and other glass vessels are graduated. The power of the acid to etch glass may be observed by having a leaden saucer, into which some powdered fluor spar and sulphuric acid are placed. The glass to be etched is first covered with a coating of wax, and the device drawn upon it with a penknife so as to remove the wax from those parts to be etched. The glass, which should cover the leaden saucer, is placed on the top, with the waxed side downwards. The saucer should be gently heated, care being taken that the temperature does not rise sufficiently to melt the wax. In a few minutes the glass may be removed and cleaned by holding it near a fire, or over a Bunsen flame, until the wax melts, when it may be wiped off, and the glass polished with a clean cloth.

The following equation explains the reaction of hydrofluoric acid on the silica of glass:—

$$SiO_2 + 4HF = SiF_4 + 2OH_2.$$

Like carbon and boron, silicon can be obtained in the three allotropic forms.

Amorphous silicon can be obtained by heating potassic silicofluoride with eight or nine times its weight of potassium, in a glass or iron tube; the product is carefully washed to free it from the potassic fluoride and excess of potassium:—

$$\underset{\substack{\text{Potassic}\\\text{silicofluoride.}}}{SiF_4,2KF} + 2K_2 = Si + \underset{\substack{\text{Potassic}\\\text{fluoride.}}}{6KF.}$$

The silicon remains as a brown amorphous powder, insoluble in water and nitric acid. It burns brilliantly in the air, becoming silicic anhydride.

Graphitoid Silicon.—This is prepared by fusing five parts of potassic silicate, ten of *cryolite*, and one of aluminium in a fireclay crucible. The button of aluminium when treated with hydrochloric acid is found to contain silicon in scaly crystals.

Diamond or Adamantine Silicon is obtained by passing the vapour of silicic chloride over aluminium:—

$$\underset{\substack{\text{Silicic}\\\text{chloride.}}}{3SiCl_4} + 2Al_2 = \underset{\substack{\text{Aluminic}\\\text{chloride.}}}{2'Al'''_2Cl_6} + Si_3.$$

The silicon is obtained in octahedra, which are sufficiently hard to scratch glass.

Silicic Anhydride (*Silica*), SiO_2, has already been stated to be an abundant natural production; its purest natural form is *rock crystal*, which occurs in six-sided prisms, terminated by six-sided pyramids. The crystals are hard enough to scratch glass; they melt to a transparent glass when heated before the oxyhydrogen blowpipe. The crystals are insoluble in water and all acids except hydrofluoric.

In its crystallized state, coloured with some metallic oxide, it forms precious stones, such as *bloodstone*, *agate*, *jasper*, *onyx*, *carnelian*, and *opal*. Silica exists abundantly in many plants. It gives stiffness to the straw of wheat, barley, and oats.

Amorphous silica is obtained by heating silicic acid to 100°:—

$$\begin{array}{ccccc} SiHo_4 & = & SiO_2 & + & 2OH_2. \\ \text{Tetrabasic} & & \text{Amorphous} & & \text{Water.} \\ \text{silicic acid.} & & \text{silica.} & & \end{array}$$

Silicic Acid,—*tetrabasic*, $SiHo_4$; *dibasic*, $SiOHo_2$.—This is a gelatinous substance, obtained by pouring hydrochloric acid into a solution of potassic or sodic silicate (soluble glass); the mixed solutions must be evaporated to some extent, and then allowed to cool:—

$$\begin{array}{ccccccc} SiKo_4 & + & 4HCl & = & SiHo_4 & + & 4KCl. \\ \text{Potassic} & & \text{Hydrochloric} & & \text{Silicic} & & \text{Potassic} \\ \text{silicate.} & & \text{acid.} & & \text{acid.} & & \text{chloride.} \end{array}$$

The disintegration of granite, which contains mercury. The jar containing the mercury is now nearly filled with water; and a gentle heat is applied to the flask. The decomposition is rather complex. The first effect is the disengagement of hydrofluoric acid, by the action of the sulphuric acid upon the calcic fluoride, and this, in contact with the silicic anhydride of the glass, undergoes decomposition, forming water and silicic fluoride, which latter escapes from the flask as a gas; but when it comes in contact with water, after it has passed through the mercury, it is decomposed, silicic and hydrofluosilicic acids being formed. Decompositions in the flask:—

$$\text{(1.)}\ CaF_2 + SO_2Ho_2 = SO_2Cao'' + 2HF.$$

$$\text{(2.)}\ 4HF + SiO_2 = 2OH_2 + SiF_4.$$

Decomposition in contact with the water, after passing through the mercury:—

$$3SiF_4 + 4OH_2 = \underset{\substack{\text{Silicic}\\ \text{acid.}}}{SiHo_4} + \underset{\substack{\text{Hydrofluosilicic}\\ \text{acid.}}}{2SiF_4,2HF.}$$

The silicic acid forms a jelly-like precipitate, while the hydrofluosilicic acid remains in solution, and may be filtered off. If a solution of a potassium salt be added to the hydrofluosilicic acid, *potassic silicofluoride*, used in the preparation of amorphous silicon, is obtained as a gelatinous mass. Barium is precipitated from solution of its salts by hydrofluosilicic acid as baric silicofluoride. A solution of the acid is kept in the laboratory for the purpose of separating barium and strontium, the salt of the former metal being insoluble in alcohol, while that of the latter is soluble.

By evaporating *in vacuo* a solution of $SiHo_4$ at 16° the dibasic acid, $SiOHo_2$, is said to be obtained.

The principal soluble silicates are potassic and sodic silicates (soluble glass); for an account of the latter, see Sodic Silicate (p. 188).

Silicic Fluoride, SiF_4.—This is a gas, and is always produced when hydrofluoric acid acts on glass or silicic anhydride :—

$$SiO_2 + 4HF = 2OH_2 + SiF_4.$$

Silicic Hydride, SiH_4.—This is a gas spontaneously inflammable in air, obtained by acting on magnesic silicide with hydrochloric acid :—

$$\begin{array}{ccccccc} Si^{iv}Mg''_2 & + & 4HCl & = & SiH_4 & + & 2MgCl_2. \\ \text{Magnesic} & & \text{Hydrochloric} & & \text{Silicic} & & \text{Magnesic} \\ \text{silicide.} & & \text{acid.} & & \text{hydride.} & & \text{chloride.} \end{array}$$

When it burns in air, it yields silicic anhydride and water :—

$$SiH_4 + 2O_2 = SiO_2 + 2OH_2.$$

Silicates form a considerable portion of the solid crust of the earth; the names and formulæ of some of the most commonly occurring ones are here given

PENTAD ELEMENTS (*continued*).

Phosphorus—P_4.

Atomic weight, 31. Sp. gr. of vapour, 62. Molecular weight, 124. Mulecular volume, normal. 1 litre of vapour weighs 62 criths. Atomicity, V. and III.

This substance was discovered by Brandt, of Hamburgh, about 200 years ago. Its name is derived from φως, *light*, and φερω, *I bear*.

Occurrence in Nature.—It is a constituent of the older rocks, in which it occurs as phosphates; by the crumbling down of these rocks it passes into the soil, from which it is taken up by plants, and it then passes as food into the bodies of animals, where in the form of calcic phosphate it serves the purpose of giving stiffness and rigidity to the animal frame. Besides forming the bony skeleton, it is found in the brain, nerves, and urine. It thus occurs in the mineral, vegetable, and animal kingdoms.

Preparation.—Brandt made his phosphours from urine, but it is now manufactured almost exclusively from *calcined bones*, or *sombrera guano*, both of which are calcic phosphate. The following is an outline of the process:—Well-calcined bones are reduced to powder, and mixed with two-thirds their weight of dilute sulphuric acid; after some time the mixture is filtered. The calcic sulphate which has been formed by the sulphuric acid is an insoluble compound, and easily separated on the filter. The liquor is then evaporated to the consistence of treacle, mixed with charcoal powder, and transferred to an iron vessel, which is heated till all moisture is expelled. After this mixture has been thoroughly dried it is heated strongly in a stoneware retort, from which a wide tube dips into water contained in the receiver. The phosphorus distils over in a soft condition, and is moulded in hot water into cylindrical sticks.

The following chemical changes take place during the above process:—Bone ashes are tricalcic diphosphate; sulphuric acid removes two of the three atoms of calcium as under:—

$$P_2O_2Cao''_3 + 2SO_2Ho_2 = P_2O_2Ho_4Cao'' + 2SO_2Cao''.$$

Tricalcic diphosphate.	Tetrahydric-calcic diphosphate.	Calcic sulphate.

By the evaporation of the liquid containing the tetrahydric-calcic diphosphate and subsequent drying with charcoal, water is expelled, and the residue is calcic metaphosphate, which when distilled with charcoal yields carbonic oxide and phosphorus, and leaves a residue of tricalcic diphosphate, thus:—

$$5P_2O_4Cao'' + C_{10} = P_2O_2Cao''_3 + 10CO + P_4.$$

Calcic metaphosphate	Tricalcic phosphate.	Carbonic oxide.

The mixture of tetrahydric-calcic diphosphate and calcic sulphate obtained by treating bones with sulwater. It readily dissolves in carbonic disulphide; if this solution be dropped upon paper, linen, &c., it quickly dries up, leaving the phosphorus in a finely divided state, which immediately bursts into flame, setting fire to the article with which it is in contact. Phosphorus takes fire spontaneously in the open air at a temperature a little above its fusing point, becoming phosphoric anhydride.

Red or *amorphous phosphorus* is an allotropic modification of phosphorus. It is obtained by heating common phosphorus for several hours in an atmosphere of carbonic anhydride at a temperature of 230° to 250°. A red powder, insoluble in carbonic disulphide, is the result: this variety of phosphorus can be kept in dry air without change; but when it is heated up to 260° it bursts into flame, passing at the same time into common or vitreous phosphorus. This variety is not poisonous, does not become luminous in the dark, has no tendency to combine with oxygen unless heated strongly, and may be carried about without fear. It was discovered by Professor Schrotter.

White, *black*, and *viscous phosphorus* are also known. The first is produced when common phosphorus is exposed under water to light for some time, and is frequently found as a crust on sticks of phosphorus which have been kept.

Phosphorus is extensively used for the manufacture of lucifer matches. As the phosphorus would be likely to ignite by exposure to the atmosphere, it is mixed up with glue or gum-water. The ends of the matches are first dipped in sulphur, paraffine, or stearine, then into the phosphorous mixture, and afterwards into a thick mixture of potassic chlorate, potassic nitrate, or some other substance rich in oxygen and gum; they are then removed to a stove and carefully dried. About 200,000 to 250,000 lbs. of phosphorus are used annually in London in the manufacture of these matches. When the surface of the match is broken by friction the phosphorus first takes fire; this inflames the sulphur, which ignites the wood: the cracking noise is owing to the potassic chlorate. *Red* phosphorus is gradually superseding the common variety for this manufacture, as its use is less liable to be attended by risk of fire or explosion; moreover, the workmen who have to breathe the fumes of common phosphorus are often afflicted with a distressing disease of the lower jaw, to which those who use red phosphorus are not liable.

Compounds of Phosphorus with Hydrogen.

Gaseous Phosphoretted Hydrogen, PH_3.

1. This gas is most readily obtained by boiling a small quantity of a strong solution of potassic or sodic hydrate, to which a small piece of phosphorus has

Fig. 14.) Half a fluid ounce of the potassic hydrate solution and ten grains of phosphorus are suitable quantities to use. The fluid should be heated gently until a flash is seen in the retort, when the beak should be immediately placed under water; the gas will quickly begin to pass over, and will ignite spontaneously as soon as it comes in contact with the atmosphere: the product of the combustion, phosphoric acid, forms beautiful revolving rings of white smoke. The experiment is a very pretty one, and is easy to perform if the above conditions are complied with. A gentle heat should be applied *continuously*, as otherwise the water will recede into the retort; for the gas is somewhat soluble in water. When the experiment is drawing to a close the beak of the retort should be removed from the water, and then the lamp may be taken away. The following is the reaction:—

$$\underset{\substack{\text{Potassic}\\\text{hydrate.}}}{3OKH} + P_4 + 3OH_2 = \underset{\substack{\text{Phosphoretted}\\\text{hydrogen.}}}{PH_3} + \underset{\substack{\text{Potassic}\\\text{hypophosphite.}}}{3POH_2Ko.}$$

2. Phosphoretted hydrogen may also be obtained by heating in a retort either phosphorous or hypophosphorous acid:—

$$a.\ \underset{\text{Phosphorous acid.}}{4POHHo_2} = \underset{\text{Phosphoric acid.}}{3POHo_3} + PH_3.$$

$$b.\ \underset{\text{Hypophosphorous acid.}}{2POH_2Ho} = \underset{\text{Phosphoric acid.}}{POHo_3} + PH_3.$$

The gas prepared by (2) is not spontaneously inflammable, but it is pure; that obtained by (1) contains free hydrogen and vapour of liquid phosphoretted hydrogen ($'P''_2H_4$), the presence of which last substance seems to be essential to its spontaneous inflammability. A solid phosphoretted hydrogen is also known.

Phosphoretted hydrogen combines directly with hydriodic and hydrobromic acids in the nascent condition, producing bodies analogous to ammonium compounds :—

$$PH_3 + HI = PH_4I.$$

Phosphoretted hydrogen.	Hydriodic acid.	Phosphoric iodide.

Phosphorous Trichloride, PCl_3.—This substance is much used in organic investigations. It is readily prepared by passing chlorine into a retort containing heated phosphorus. The phosphorus must be kept in excess. The experiment requires care, or it may prove dangerous. Phosphorous chloride is a colourless liquid, with an offensive odour. Water decomposes it into hydrochloric and phosphorous acids :—

$$PCl_3 + 3OH_2 = 3HCl + POHHo_2.$$

Phosphorous chloride.	Water.	Hydrochloric acid.	Phosphorous acid.

Phosphoric Chloride, PCl_5.—This substance is also much used in organic chemistry. It is prepared by passing chlorine through the trichloride. It is a

The chief differences are these:—Phosphorus does not form salts—such as AmCl, SO_2Amo_2, &c.—which would have the composition PH_4Cl, $SO_2(PH_4O_2)$, &c., with the exception of $(PH_4)'I$ and $(PH_4)'Br$, already noticed (p. 124); yet in organic chemistry even this difference disappears, and salts of these types are well known, containing phosphorus in the place of nitrogen.

Phosphorus forms more acids than nitrogen. Nitric anhydride assimilates only one molecule of water, thus:—$N_2O_5 + OH_2 = 2NO_2Ho$; whereas phosphoric anhydride takes three, thus:—$P_2O_5 + 3OH_2 = 2POHo_3$; still this is only after some time, for when phosphoric anhydride is added to water, only one atom of water unites at first,—$_2PO_5 + OH_2 = 2PO_2Ho$, forming an acid strictly analogous to nitric.

Oxides and Oxy-acids of Phosphorus.

Phosphorous Anhydride, P'''_2O_3, and **Phosphorous Acid,** P^vOHHo_2.—When phosphorus is slowly oxidized in a limited supply of dry air, P_2O_3 is formed, which is a white inflammable powder, which readily dissolves in water, producing the acid $POHHo_2$:—

$$\underset{\substack{\text{Phosphorous}\\\text{anhydride.}}}{P_2O_3} + 3OH_2 = \underset{\substack{\text{Phosphorous}\\\text{acid.}}}{2POHHo_2}.$$

Take a glass tube, about one foot long and half an inch in diameter; at about two inches from one end bend the tube at an obtuse angle, drawing the end of the shorter limb out into a conical shape, leaving an aperture about the size of a pin's point. Into the bent end of this tube, at the angle, place a small piece of phosphorus, carefully dried by folding it in blotting-paper. Gently heat the tube where the phosphorus lies by means of a spirit-lamp, when the phosphorus will take fire and burn with a greenish flame. Only a small quantity of air enters the tube through the small aperture. The anhydride condenses as a white powder in the upper cold part of the tube. This is solid phosphorous anhydride.

Phosphorous acid may also be prepared from phosphorous iodide, PI_3. To prepare this phosphorous iodide, place a small piece of dry phosphorus in a test-tube, cover it with about eleven times its weight of iodine. The two bodies combine energetically without any application of heat. After the action is over, water is added, when the following decomposition takes place:—

$$PI_3 + 3OH_2 = \underset{\text{Phosphorous acid}}{POHHo_2} + \underset{\text{Hydriodic acid.}}{3HI.}$$

Phosphoric Anhydride, $P^v_2O_5$, is obtained by burning fragments of phosphorus in a large globe or jar, through which dry air is made to pass between each combustion of phosphorus. It is a snow-white

The metaphosphates are prepared by expelling by heat, water or a volatile base, or both, from an orthophosphate. The following are examples :—

$$\underset{\text{Microcosmic salt.}}{\mathrm{POHoNaoAmo}} = \underset{\substack{\text{Sodic}\\\text{metaphosphate.}}}{\mathrm{PO_2Nao}} + \underset{\substack{\text{Ammonic}\\\text{hydrate.}}}{\mathrm{AmHo}}.$$

$$\underset{\substack{\text{Dihydric-sodic}\\\text{phosphate.}}}{\mathrm{POHo_2Nao}} = \underset{\substack{\text{Sodic}\\\text{metaphosphate.}}}{\mathrm{PO_2Nao}} + \mathrm{OH_2}.$$

Metaphosphoric acid is monobasic.

Pyrophosphoric Acid, $P_2O_3Ho_4$.—This acid is tetrabasic, and is so named because its salts are obtained by heating to redness phosphates which contain two atoms of a fixed base :—

$$\underset{\substack{\text{Hydric-disodic}\\\text{phosphate.}}}{2\mathrm{POHoNao_2}} = \underset{\substack{\text{Sodic}\\\text{pyrophosphate.}}}{\mathrm{P_2O_3Nao_4}}. + \underset{\text{Water.}}{\mathrm{OH_2}}.$$

To prepare this acid, a solution of sodic pyrophosphate is added to a solution of plumbic acetate (sugar of lead), and the precipitated pyrophosphate is filtered off, washed on the filter, and then suspended in water, whilst a current of sulphuretted hydrogen is passed through the fluid. The plumbic sulphide formed is filtered off, and the fluid is a solution of pyrophosphoric acid :—

$$\underset{\substack{\text{Plumbic}\\\text{pyrophosphate.}}}{\mathrm{P_2O_3Pbo''_2}} + \underset{\substack{\text{Sulphuretted}\\\text{hydrogen.}}}{2\mathrm{SH_2}} = \underset{\substack{\text{Plumbic}\\\text{sulphide.}}}{2\mathrm{PbS}} + \underset{\substack{\text{Pyrophosphoric}\\\text{acid.}}}{\mathrm{P_2O_3Ho_4}}.$$

Phosphoric Acid, Orthophosphoric Acid, $POHo_3$.—There are several methods of preparing this acid, some of which have already been noticed (see pp. 154, 155). The phosphoric acid of the Pharmacopœia is prepared by boiling phosphorus with nitric acid in a retort, until it dissolves, and then evaporating the liquid in a porcelain dish until red fumes cease to come off. It can also be prepared by treating *bone ash* with a considerable excess of sulphuric acid; after heating, the mixture is filtered, to separate calcic sulphate, and ammonic carbonate, $COAmo_2$, is added to the filtrate till effervescence ceases. By again filtering, evaporating the filtrate to dryness, and igniting the residue to expel the ammonic salts, phosphoric acid is obtained.

Phosphoric acid is not poisonous; it is, however, a powerful acid, reddens vegetable blues, and with bases forms a large and important class of salts, called phosphates. The acid prepared with nitric acid, when boiled and then allowed to cool, forms a transparent solid resembling ice, which is hence called *glacial* phosphoric acid.

Tests.—The varieties of phosphoric acid are distinguished from one another by means of a solution of argentic nitrate, with which—

a. Orthophosphoric acid gives a canary-yellow pre-

Arsenic—As_4.

Atomic weight, 75. Molecular weight, 300. Molecular volume, normal. 1 litre of arsenic vapour weighs 150 criths. Atomicity, V and III.

Occurrence in Nature.—It is sometimes found native, but more frequently in combination with oxygen, sulphur, iron, nickel, and cobalt.

Manufacture.—The ores containing arsenic are heated in a reverberatory furnace, and the volatile products condensed in a chimney of peculiar construction. The crude arsenious anhydride thus produced is purified by sublimation. It is then reduced with charcoal in an earthenware retort, and the metal easily sublimes.

Properties.—This metal is of a steel-grey colour. When heated, it oxidizes into arsenious anhydride. As_2O_3. It combines with the metals, like sulphur and phosphorus, and unites with oxygen in two proportions, forming arsenious anhydride, As_2O_3, and arsenic anhydride, As_2O_5. Arsenic volatilizes at 180°, giving off a vapour which has a strong odour of garlic. The metal has a specific gravity of 5·7. It is crystalline and very brittle. Its molecule, like that of phosphorus, contains four atoms. No oxy-salts are known in which arsenic performs the functions of a basylous element.

Its chemical characters particularly resemble those of phosphorus, and it may be conveniently studied in connection with that substance, along with the other two pentad metals, bismuth and antimony, which also have a strong analogy to phosphorus and nitrogen. Their principal compounds are constituted on the types of the following general formulæ, in which M represents the pentad element :—

$$MH_3,\ MH_4Cl;\ M_2O_3,\ M_2O_5.$$

Principal Compounds of Arsenic.

Arseniuretted Hydrogen, AsH_3, is produced when nascent hydrogen acts on an arsenic compound: it is usually prepared by putting arsenious acid into a hydrogen apparatus along with zinc and dilute sulphuric acid :—

$$AsHo_3 + 3Zn + 3SO_2Ho_2 = 3SO_2Zno'' + 3OH_2 + AsH_3.$$

It is also readily obtained by acting on an alloy of zinc and arsenic with dilute sulphuric acid.

Arseniuretted hydrogen is a colourless gas, having the smell of garlic. It burns with a blue flame, producing arsenious anhydride and water with free access of air,—

$$2AsH_3 + 3O_2 = As_2O_3 + 3OH_2;$$

but when the supply of air is limited, as it is when a cold piece of porcelain is held in the flame, the

As_2O_3, is prepared by heating arsenic in contact with the air. It is soluble in hot water, and crystallizes in transparent octahedra. The alkalies dissolve this compound readily, forming arsenites. It is a powerful poison. In experimenting with arsenic compounds the student should avoid inhaling the vapours.

Arsenious acid, $AsHo_3$, is only known in solution. *Scheele's green* is a hydric-cupric arsenite, $AsHoCuo''$.

Arsenic Anhydride, As_2O_5.—This compound is produced by dissolving arsenious anhydride in hot hydrochloric acid, and oxidizing with nitric acid so long as red vapours are produced. The solution is then carefully evaporated to dryness. It slowly dissolves in water, the solution being slightly acid; on being evaporated, it yields, after some time, crystals of arsenic acid, $AsOHo_3$. When the anhydride is heated it is decomposed into arsenious anhydride and oxygen.

Arsenic acid resembles in the closest manner phosphoric acid:—

Arsenic acid	$AsOHo_3$.
Metarsenic acid.	AsO_2Ho.
Pyrarsenic acid	$As_2O_3Ho_4$.

Hydric-disodic arsenate, $AsOHoNao_2$, cannot be distinguished in appearance from the corresponding phosphate. It may be prepared by adding a solution of sodic carbonate to a solution of arsenic acid until the acid is completely neutralized, and then evaporating.

Diarsenious Disulphide (Realgar), As_2S_2, which is found native, and can be formed artificially by heating arsenic with the atomic proportion of sulphur, is employed in painting, and for making *white fire.*

Arsenious Sulphide (Orpiment), As_2S_3, is also a natural compound. It may be prepared by precipitating a solution of arsenious anhydride in hydrochloric acid with hydrosulphuric acid: after the lapse of some time a higher sulphur compound is precipitated.

Sulpharsenic Anhydride, As_2S_5, is not important.

The three sulphides of arsenic are sulphur anhydrides. Thus, if sulphuretted hydrogen be passed through a solution of arsenious anhydride, a precipitate of arsenious sulphide is formed. The addition of ammonic or potassic sulphide completely dissolves it, on account of the formation of ammonic or potassic sulpharsenite, thus:—

$$As_2S_3 + 3SK_2 = 2AsKs_3.$$

The addition of hydrochloric acid reproduces the precipitate, thus:—

$$AsKs_3 + 3HCL = 3KCl + AsHs_3.$$

This latter compound, being unstable, decomposes thus:—

$$\underset{\text{Sulpharsenious acid.}}{2AsHs_3} = \underset{\text{Sulphuretted hydrogen.}}{3SH_2} + \underset{\text{Sulpharsenious anhydride.}}{As_2S_3}$$

precipitate of arsenious sulphide is thrown down, soluble in ammonic hydrate, and can be re-precipitated by sulphuric or hydrochloric acid.

The detection of arsenious acid when mixed with the contents of the stomach is much more difficult. The presence of organic matter so interferes with the action of the tests just given as to render their results worthless in a matter of so much importance. In investigations of this kind the two following tests are applied.

Marsh's Test.—This test depends on the fact that nascent hydrogen combines with arsenic, as already explained, and forms therewith arseniuretted hydrogen. A hydrogen bottle, fitted with a tube bent at right angles, and terminating in a jet, is supplied with zinc and dilute sulphuric acid, both of which should be *free from arsenic.* When the air has been all chased out a light may be applied to the jet, the flame of which should not stain a surface of white porcelain when brought in contact with it. When this condition is fulfilled, the substance to be tested may be introduced into the hydrogen bottle, and after a short time the jet of gas relighted and tested again with the white porcelain; if arsenic be present a brilliant black metallic film will appear on the porcelain. If the delivery tube be heated to a red heat the gas will be decomposed, and arsenic deposited as a metallic ring inside the tube, behind the heated portion. The gas passed through solution of argentic nitrate precipitates metallic silver as a black powder.

Reinsch's Test.—The poisoned liquid is made acid with hydrochloric acid and slips of clean copper foil, wire, or gauze introduced. The liquid is now boiled for a few minutes, when, if arsenic be present, it will be found on the surface of the copper as a grey film.

The slips of copper are carefully washed, dried, introduced into a glass tube, which is then heated; the arsenic is quickly converted into arsenious anhydride, which condenses on the cold surface of the tube in minute octahedral crystals.

Antimony—Sb.

Atomic weight, 122. Probable molecular weight, 488. Atomicity, V. and III.

Occurrence in Nature.—Chiefly in combination with sulphur as grey antimony ore, Sb_2S_3, called also *stibnite*; also as an oxide.

Preparation.—The ore is first freed from earthy impurities by fusion; the sulphur is then removed from it by heating it with metallic iron or potassic carbonate:—

$$\underset{\text{Antimonious sulphide.}}{Sb_2S_3} + 3Fe = \underset{\text{Antimony.}}{Sb_2} + \underset{\text{Ferrous sulphide.}}{3FeS}$$

Antimony is also obtained by roasting the ore, by which treatment it is partially converted into oxide; this is then reduced by fusing it with charcoal and *metal* of the printer; and with tin, *Britannia metal.* It is too brittle to be used alone, but its alloys have numerous applications.

Principal Compounds.

Antimoniuretted Hydrogen, SbH_3, also called *Stibine.*—This compound is produced when a soluble antimony compound is acted on by nascent hydrogen, and is obtained by the same process as that which yields arseniuretted hydrogen (*arsine*), which it closely resembles. It burns with a greenish flame, forming antimonious oxide and water. A cold piece of porcelain held in the flame is stained by a film of metallic antimony, and antimony is also deposited as a metallic ring when the tube through which the gas passes is heated to redness. When passed through a solution of argentic nitrate, a black precipitate, antimonious argentide, $SbAg_3$, is formed; this property serves to distinguish it from arseniuretted hydrogen.

Antimonious Chloride, $SbCl_3$, *butter of antimony.*—This compound is obtained when hydrosulphuric acid is generated by the action of hydrochloric acid on the native antimonious sulphide:—

$$Sb_2S_3 + 6HCl = 3SH_2 + 2SbCl_3.$$

The residual solution is then distilled until each drop of the condensed vapour, on falling into the water of the receiver, produces a white precipitate. The receiver is then changed, and the process carried on. The air should be carefully excluded. Alkaline solutions decompose antimonious chloride, producing the corresponding hydrate. Powdered antimony takes fire in chlorine gas, forming this chloride. With water it undergoes decomposition, forming antimonious oxy-chloride as a white precipitate:—

$$SbCl_3 + OH_2 = 2HCl + SbOCl.$$

Antimonious chloride. — Antimonious oxychloride.

Antimonic Chlorine, $SbCl_5$, is prepared by passing a stream of dry chlorine over metallic antimony, gently heated in a porcelain tube, or better still, by passing chlorine over antimonious chloride. It is a colourless volatile liquid, which produces antimonic oxychloride, $SbOCl_3$, when acted on with a small quantity of water, and antimonic acid, $SbOHo_3$, when an excess of water is used. It will be noticed that the constitution of this acid resembles that of phosphoric acid, $POHo_3$.

Antimonious Sulphide, Sb_2S_3, is the native sulphide from which the metal is obtained. It may be prepared artificially by passing sulphuretted hydrogen into solution of antimonious chloride. The sulphide separates as a fine orange-coloured precipitate :—

$$2SbCl_3 + 3SH_2 = Sb_2S_3 + 6HCl.$$

Antimonious chloride. — Antimonious sulphide.

Antimonic Anhydride, Sb_2O_5, is produced by heating any of its acid derivatives, which are the exact analogues of the phosphoric acids:—

a. Orthantimonic acid . .	$SbOHo_3$.
b. Metantimonic „ . .	SbO_2Ho.
c. Pyrantimonic „ . .	$Sb_2O_3Ho_4$.

a is obtained by the action of water on antimonic chloride.

b. *Aqua regia* converts metallic antimony into this variety.

c is formed when hydrochloric acid is added to solutions of its salts.

Tests.—Sulphuretted hydrogen precipitates antimony from its solutions as an orange-coloured sulphide, which is soluble in yellow ammonic sulphide or sodic hydrate.

Antimony gives a metallic film on porcelain with *Marsh's test*, and when the gas is passed into solution of argentic nitrate the metal is precipitated as antimonious argentide, which is soluble in tartaric acid. Arsenic, under like conditions, remains in solution.

Zinc precipitates antimony from its solutions as a black powder.

Bismuth—Bi.

Atomic weight, 208. Atomicity, V. and III.

Occurrence in Nature.—Chiefly in the metallic state mixed up in various rocks.

Preparation.—It is separated from its earthy impurities by fusion.

Properties.—This metal is of a reddish white colour. It is highly crystalline, and very brittle; its sp. gr. is 9·83; it melts at about 265°, and volatilizes at a high temperature. It is the most diamagnetic of and organic substances. It is little oxidized by air in the cold, but burns with a blue flame when strongly heated. Its best solvent is nitric acid. Its compounds are analogous to those of antimony, except that it forms no compound with hydrogen.

Fusible metal is an alloy of 8 parts bismuth, 5 of lead and 3 of tin. It melts below 100°. Bismuth and antimony form the elements in the construction of the thermo-electric pile.

Principal Compounds.

Bismuthous Chloride, $BiCl_3$, is produced by processes similar to those which yield the corresponding antimony compound, to which the student is referred. Its reactions with water are also analogous.

Bismuthous Oxide, Bi_2O_3.—This oxide is obtained by heating bismuthous nitrate, or by heating the precipitate formed by adding potassic hydrate to bismuthous chloride, $BiCl_3$. It fuses at a high temperature, and forms a powerful flux with silicious substances.

Bismuthic Oxide or Anhydride, Bi_2O_5, is made by heating bismuthic acid. It is readily reduced to

Bismuthous Nitrate (trisnitrate of bismuth), N_3O_6Bio''',$5OH_2$.—This salt is prepared by dissolving bismuth in nitric acid. If the solution be left for some time, large crystals of bismuthous nitrate are deposited. A basic nitrate is produced when a solution of bismuthous nitrate is treated with a large quantity of water. This salt was extensively used as a cosmetic, but it injures the skin. Bismuthous nitrate is occasionally used in medicine,

Tests.—Water added in large quantities to solutions of bismuth salts causes the formation of basic salts as white precipitates.

With sulphuretted hydrogen a black sulphide is obtained, insoluble in ammonic sulphide.

MONAD METALS.

Potassium—K_2.

Atomic weight, 39. Molecular weight (probable), 78.
Atomicity, 1. Sp. gr., 0·865.

Special Notice.—Potassium was discovered by Sir Humphry Davy, in 1807, by decomposing moist potassic hydrate (KHo) by means of a powerful galvanic battery. Its discovery has been of great importance to the progress of chemical science.

Occurrence in Nature.—Potassium occurs in nature as silicate, $SiKo_4$, in combination with aluminic silicate, in *granite*, *felspar*, *mica*, and other igneous rocks.

These rocks are decomposed by the action of the carbonic anhydride and moisture of the atmosphere, potassic carbonate ($COKo_2$) being formed (see p. 146).

The potassic carbonate, being soluble in water, gradually dissolves, and so passes to the soil, from which it is taken up by plants, and the potassium is found in their juices, generally in the form of an organic acid salt.[*] It occurs also as potassic chloride (KCl) in saline deposits, and as chloride, bromide, iodide, and sulphate, in sea and mineral waters.

Preparation.—Potassium is prepared by the following process:—A mixture of potassic carbonate and charcoal is prepared by heating hydric-potassic tartrate (cream of tartar). When cold the mass is reduced to powder and mixed with 1-10th of its weight of charcoal in small lumps. It is then transferred to a well-hammered iron retort, or a bottle in which mercury is imported. The retort is placed in a furnace and the fire supplied with coke and charcoal, so that the retort is enveloped in flame and kept at a uniform heat, approaching to whiteness. A copper receiver, divided by a partition, into one division of which a tolerably large pipe from the retort dips into naphtha and the other division filled with ice, is connected with the retort. The fire is now raised in temperature, and decomposition of the potassic carwith oxygen, and then into potassic hydrate (KHo) by double decomposition with the moisture; consequently it has to be preserved beneath rock oil or naphtha. Thrown upon the surface of cold water, it decomposes it with such energy as to cause the liberated hydrogen to burn; the resulting compound being potassic hydrate, which dissolves in the water:—

$$2OH_2 + K_2 = \underset{\text{Potassic hydrate.}}{2KHo} + H_2.$$

When heated in dry air it takes fire, burning with a purple flame, and forming potassic oxide (OK_2), thus:—

$$2K_2 + O_2 = 2OK_2.$$

Uses.—Owing to its powerful attraction for oxygen and the other non-metallic elements, potassium is of use in the preparation of several of the elements, as boron, silicon, and aluminium (see pp. 85, 143).

Principal Compounds.

Potassic Oxide (potash), OK_2.—This substance is prepared, as already stated, by exposing potassium to the action of dry oxygen, or by heating potassic hydrate with potassium:—

$$2KHo + K_2 = 2OK_2 + H_2.$$

It is a white deliquescent solid, which is decomposed by water, becoming incandescent, and forming potassic hydrate:—

$$OK_2 + OH_2 = 2KHo.$$

Potassic Hydrate (caustic potash), KHo.—This important compound is prepared (1) by decomposing potassic carbonate by calcic hydrate ($CaHo_2$), thus:—

$$\underset{\substack{\text{Potassic}\\\text{carbonate.}}}{COKo_2} + \underset{\substack{\text{Calcic}\\\text{hydrate.}}}{CaHo_2} = \underset{\substack{\text{Potassic}\\\text{hydrate.}}}{2KHo} + \underset{\substack{\text{Calcic}\\\text{carbonate.}}}{COCao''}.$$

The calcic carbonate ($COCao''$), being insoluble, precipitates, and the potassic hydrate remains in solution; the solution is decanted off, evaporated to dryness in an iron or silver dish, and the heat continued till the mass fuses, when it is run into cylindrical moulds; when cold the sticks thus formed are removed to well-stoppered bottles, as otherwise it would absorb and combine with the moisture and carbonic anhydride contained in the atmosphere. (2) By the action of potassium upon water, as already mentioned:—

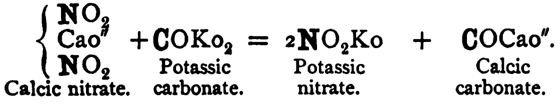
$$2OH_2 + K_2 = 2KHo + H_2.$$

Potassic hydrate, when fused, is a hard, greyish white solid, having a sp. gr. of 2·2. It can be obtained in crystals forming acute rhombohedra, and having the composition $KHo, 2OH_2$, by making a hot concentrated solution in water. It is highly alkaline, and neutralizes the most powerful acids. By dissolving it in water heat is evolved. It is used in the manufacture of soft soap, soluble glass (potassic silicate.) and in medicine as a caustic for dissolving the skin. In the laboratory it is an indispensable reagent, being ashes are treated with water, in which the potassic carbonate readily dissolves, and after allowing the insoluble portions to settle, the solution is decanted off and evaporated to dryness. The pure carbonate is obtained by heating hydric-potassic tartrate (cream of tartar), dissolving out the resulting carbonate, filtering from the charcoal with which it is mixed, and evaporating to dryness.

Potassic carbonate crystallizes in oblique rhombic octahedral crystals, which are deliquescent. Its solution has an alkaline taste and reaction. It is used in the manufactures of soap and glass, and in the preparation of most other compounds of potassium.

Hydric-potassic Carbonate (bicarbonate of potash), $\mathbf{C}OHoKo$.—When carbonic anhydride is passed into a solution of potassic carbonate, a white crystalline precipitate of the hydric-potassic carbonate is produced :—$\mathbf{C}OKo_2 + \mathbf{O}H_2 + \mathbf{C}O_2 = 2\mathbf{C}OHoKo$. This is the process employed for preparing the salt. It is less soluble in water than potassic carbonate, and has no alkaline action on litmus. Owing to its great purity if well crystallized, it is employed for the preparation of other potassium compounds. It is also used in medicine, and in making effervescing draughts.

Potassic Nitrate (nitre, saltpetre), $\mathbf{N}O_2Ko$.—This salt is brought to this country chiefly from the East Indies, where it appears, in some parts, like snow on the surface of the ground. The nitre is collected, dissolved in water, and a little potassic carbonate added to decompose the calcic nitrate which it usually contains, by precipitating the calcium as carbonate :—



$$\underset{\text{Calcic nitrate.}}{\left\{\begin{matrix}\mathbf{N}O_2\\ Cao''\\ \mathbf{N}O_2\end{matrix}\right.} + \underset{\text{Potassic carbonate.}}{\mathbf{C}OKo_2} = \underset{\text{Potassic nitrate.}}{2\mathbf{N}O_2Ko} + \underset{\text{Calcic carbonate.}}{\mathbf{C}OCao''}.$$

The liquid is evaporated, and the nitre of commerce crystallizes out. In France large quantities of it are manufactured in the following manner:—Heaps of old mortar and earth are collected in open sheds, and freely exposed to the air. These heaps are constantly turned and watered with putrid urine. This process, which is sometimes called *nitrification*, causes the nitrogenous matters (as ammoniacal salts, urea, &c.) contained in the urine to become converted into nitric acid, which, acting on the calcic oxide of the mortar, forms calcic nitrate, $\left\{\begin{matrix}NO_2\\NO_2\end{matrix}\right.Cao''$. When a large quantity of this salt has been formed, the mortar heaps are treated with water in tanks; this dissolves out the calcic nitrate; to the solution, after separating it from the insoluble matter, a solution of crude potassic carbonate is added: this causes a double decomposition, potassic nitrate being formed and remaining in solution, while calcic carbonate is precipitated:—

oxygen, a deliquescent mass of potassic nitrate ($\mathbf{N}OKo$) being formed :—

$$\underset{\text{Potassic nitrate.}}{2\mathbf{N}O_2Ko} = \underset{\text{Potassic nitrite.}}{2\mathbf{N}OKo} + O_2.$$

It is used as an oxidizing agent, both in the laboratory and manufactory,* and very largely in the manufacture of gunpowder, Gunpowder is a mixture of potassic nitrate, charcoal, and sulphur, in the following proportions by weight :—

Potassic nitrate . . .	75 parts.
Charcoal (carbon) . .	15 „
Sulphur	10 „
	100 „

The best charcoal for this purpose is that prepared from dogwood or elder. When the gunpowder is exploded, a quantity of gas, principally nitrogen and carbonic anhydride, is generated, which, owing to the intense heat to which it is subjected, occupies 1,500 times the bulk of the gunpowder used, thus accounting for the violence of the explosion. Generally the reaction may be said to be the following, though the actual results are by no means so simple :—

$$\underset{\substack{\text{Potassic}\\ \text{nitrate.}}}{4\mathbf{N}O_2Ko} + 6C + S_2 = 6CO_2 + 2N_2 + \underset{\substack{\text{Potassic}\\ \text{sulphide.}}}{2SK_2}.$$

Potassic Sulphate, SO_2Ko_2.—This salt is the residue of the process for manufacturing nitric acid from potassic nitrate ; thus :—

$$\underset{\substack{\text{Potassic}\\ \text{nitrate.}}}{2\mathbf{N}O_2Ko} + \underset{\substack{\text{Sulphuric}\\ \text{acid.}}}{SO_2Ho_2} = \underset{\substack{\text{Potassic}\\ \text{sulphate.}}}{SO_2Ko_2} + \underset{\substack{\text{Nitric}\\ \text{acid.}}}{2\mathbf{N}O_2Ho}.$$

---

* Also as an *antiseptic*, that is, as a preventive of putrefaction in meat, &c.,—occasionally in making nitric acid.

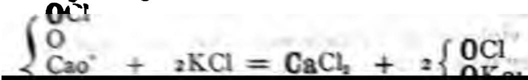
It crystallizes in two distinct crystalline forms, and hence is *dimorphous*. It is not very soluble in water. It is used in the manufacture of *potash-alum*.

Hydric-potassic Sulphate, SO_2HoKo.—When twice as much sulphuric acid is used as in the above case, this salt is the result:—

$$2NO_2Ko + 2SO_2Ho_2 = 2SO_2HoKo + 2NO_2Ho.$$

It crystallizes in rhomboidal tables. It is used occasionally as a flux, and also in calico-printing.

Potassic Chlorate, $\left\{\begin{matrix}OCl\\OKo\end{matrix}\right.$.—This salt is prepared by passing chlorine gas through a mixture of calcic hydrate ($CaHo_2$) and potassic chloride (KCl). Calcic hydrochlorite ($CaOCl_2$) is first formed, and as the temperature rises it is decomposed, calcic chlorate being formed, which undergoes mutual exchange with the potassic chloride: potassic chlorate and calcic chloride ($CaCl_2$) being formed, thus (see pp. 81, 83):—



$$\left\{\begin{matrix}OCl\\O\\Cao\end{matrix}\right. + 2KCl = CaCl_2 + 2\left\{\begin{matrix}OCl\\ \end{matrix}\right.$$

It deflagrates violently with combustible matter, explosions occurring with very slight blows. It is largely used in the manufacture of lucifer matches; as an oxidizing agent in calico-printing; and when added to a solution acidulated with hydrochloric acid, is often used in the laboratory for the same purpose. It is also used in the mixture with which the friction tubes used for firing cannon are charged.

Potassic Chloride, KCl.—This salt occurs in salt deposits at Stassfurt, in Germany, and other places. It is also obtained from *kelp*, the ashes of burnt seaweed. It crystallizes in cubes, is readily soluble in water, and is used in the manufacture of *potash-alum* and potassic chlorate.

Potassic Iodide, KI.—This salt occurs to a slight extent in sea-water. It is prepared by digesting iodine and iron filings together with water till the solution is colourless. Ferrous iodide (FeI_2) is thus formed; the filtered solution is mixed with a solution of pure potassic carbonate, which causes ferrous carbonate ($COFeo''$) to precipitate, while potassic iodide remains in solution:—

$$\underset{\substack{\text{Ferrous}\\ \text{iodide.}}}{FeI_2} + \underset{\substack{\text{Potassic}\\ \text{carbonate.}}}{COKo_2} = \underset{\substack{\text{Potassic}\\ \text{iodide.}}}{2KI} + \underset{\substack{\text{Ferrous}\\ \text{carbonate.}}}{COFeo''}.$$

The solution is filtered to separate the suspended precipitate, and evaporated till the salt crystallizes on cooling.

Potassic iodide crystallizes in cubes, and is very soluble in water: it is also soluble in alcohol. Its solution dissolves iodine, forming a deep brown liquid. It is used in photography and in medicine.

Tests for Potassium Compounds.—If tartaric acid be added to a solution of a potassium salt in a test-tube, the mixture agitated thoroughly and allowed to stand for a short time, a white crystalline precipitate of hydric-potassic tartrate (cream of tartar) is produced.

Platinic chloride ($PtCl_4$) gives in solutions of potassium salts, to which a little hydrochloric acid has been added, a yellow crystalline precipitate after agitation, which is a molecular combination of potassic chloride and platinic chloride ($2KCl,PtCl_4$).

Compounds of potassium, when introduced on a loop of platinum wire into the Bunsen or blowpipe flame, tinge it of a purple or violet colour; and if the flame be examined by the spectroscope, the spectrum will be found to consist of three bright lines, two in the red portion and one in the violet.

Sodium—Na.

Atomic weight, 23. Molecular weight (probable), 46.
Atomicity, I. Sp. gr., 0·97.

Special Notes.—Sodium was discovered by Sir Humphry Davey in 1807, immediately after he dis-

($CONao_2$) by means of carbon. An intimate mixture of 7 parts of dried sodic carbonate, 2 parts of powdered charcoal, and 1 part of finely powdered calcic carbonate is made into a paste with oil, and after carbonizing the oil by heating the mixture in an iron pot, it is transferred to an iron retort like that used for preparing potassium, and is distilled. The reaction is the same as occurs in preparing potassium; the calcic carbonate is used simply to prevent the charcoal separating from the sodic carbonate when the latter fuses :—

$$\underset{\text{Sodic carbonate.}}{CONao_2} + 2C = Na_2 + \underset{\text{Carbonic oxide.}}{3CO}.$$

The process is more manageable than that for preparing potassium.

Other processes by which it may be obtained are—(1) by electrolysis of sodic hydrate (Davy's original method); (2) by decomposing sodic hydrate at a white heat by means of metallic iron :—

$$\underset{\text{Sodic hydrate.}}{4NaHo} + 3Fe = 2Na_2 + \underset{\substack{\text{Triferric}\\\text{tetroxide.}}}{Fe_3O_4} + 2H_2.$$

Properties.—Sodium is a silver-white metal, soft at ordinary temperatures. It melts at 90°, and oxidizes in dry air, passing into sodic oxide (ONa_2). It boils at a red heat. Thrown upon water, sodium decomposes it with great energy, the temperature occasionally rising high enough to inflame the liberated hydrogen. In hot water the hydrogen takes fire at once, burning with a yellow flame, due to a small quantity of sodium vapour. In both cases sodic hydrate (NaHo) is the product; it dissolves in the water :—

$$2OH_2 + Na_2 = 2NaHo + H_2.$$

Uses.—Sodium is extensively used in the preparation [illegible] metals magnesium and aluminium, and [illegible] found advantageous by Mr. Crookes [illegible] sodium to the mercury used in the [illegible] from its ores. Being cheaper than [illegible] weight being less, and being [illegible] power of action, it may be advantageously [illegible] substitute for that metal in all possible [illegible] it is used as a reducing agent; [illegible] sodic hydrate (NaHo); and many [illegible]

Sodic [illegible] ONa_2.—As already stated, [illegible] when sodium oxidizes in dry [illegible] $= 2ONa_2$. It may likewise be prepared [illegible] sodic hydrate with sodium, thus:— [illegible] $= 2ONa_2 + H_2$. It is a yellowish [illegible] absorbs moisture from the air, [illegible] $ONa_2 + OH_2 = 2NaHo$.

Sodic Hydrate, caustic soda, NaHo.—This im[illegible] prepared by boiling a solution of [illegible] with calcic hydrate ($CaHo_2$) in an iron pan; calcic carbonate ($COCao''$) is precipiare obtained. It is extensively used in the manufacture of hard soaps and of sodic silicate (soluble glass), and is an indispensable reagent in the laboratory.

Sodic Chloride (common salt), $NaCl$.—This well-known substance occurs in nature, forming immense beds of *rock salt* in Cheshire; Durham; in Ireland, near Belfast; at Wieliczka in Poland, Cordona in Spain, Galicia, and Tyrol. It also occurs in sea water to the extent of four ounces to the gallon. A large quantity of salt is obtained from the brine springs of Droitwich; but our chief supply is from the mines of Cheshire. Rock salt is always too impure for use; and if no natural brine spring exists, one is formed by sinking a shaft into the rock salt, and introducing water. This water, when saturated, is pumped up, and evaporated in large iron pans. As the salt separates, it is removed by means of a wooden scoop, and pressed, while moist, into moulds, and then transferred to the drying stove. When large crystals are required, as for *bay salt*, the evaporation is conducted very slowly. The commercial article generally contains a little magnesic chloride, which renders it somewhat deliquescent. The following analysis shows the total *salts* in an imperial gallon of Droitwich brine:—

Sodic chloride . . .	21,761·872	grains.
Magnesic chloride ($MgCl_2$)	2·560	,,
Calcic sulphate (SO_2CaO) .	91·120	,,
Aluminic sulphate . .	14·400	,,
Sodic sulphate (SO_2NaO_2)	342·720	,,
Sodic iodide . . .	·208	grain.
	22,212·880	grains.

It is also obtained from sea water by evaporation. Sodic chloride crystallizes in cubes, and is almost equally soluble in cold and hot water. It is extensively used in the manufacture of sodic carbonate, in glazing stoneware, as an antiseptic in preserving meat, and it is an essential constituent of animal food.

Sodic Sulphate (Glauber's salt), SO_2Nao_2; crystallized, $SO_2Nao_2 10OH_2$.—This compound has been found pure in nature, forming the mineral called *Thénardite.* It is prepared in enormous quantities in *Leblanc's process* for making sodic carbonate, by acting upon sodic chloride with sulphuric acid in an iron pan heated by a furnace:—

$$\underset{\text{Sodic chloride.}}{2NaCl} + \underset{\text{Sulphuric acid.}}{SO_2Ho_2} = \underset{\text{Sodic sulphate.}}{SO_2Nao_2} + 2HCl.$$

The hydrochloric acid (HCl) produced passes off in the gaseous form, and is absorbed by water trickling down large towers filled with coke. When the decomposition is complete the sodic sulphate is raked out and allowed to solidify; the solid mass is termed *salt-cake.*

Sodic sulphate is very soluble in water, and it may be obtained crystallized in four-sided prisms; the crystals may be either SO_2Nao_2; $SO_2Nao_2,7OH_2$; or $SO_2Nao_2,10OH_2$, according to the manner in which they are obtained. The varieties containing water are remarkably efflorescent.* The crystals dissolve in hydrochloric acid, with great absorption of heat from surrounding bodies, and this mixture is consequently used in chemical operations as a freezing mixture where a low temperature is required. It has a bitter taste, and is occasionally used in medicine; it is also used in making glass.

Hydric-sodic Sulphate (bisulphate of soda), SO_2HoNao.—When nitric acid (NO_2Ho) is prepared in a glass retort from sodic nitrate (NO_2Nao) and sulphuric acid (SO_2Ho_2) the residue in the retort is hydric-sodic sulphate, thus:—

---

* Becoming powdery from loss of water of crystallization.

$$\underset{\text{Sodic Nitrate.}}{NO_2Nao} + \underset{\text{Sulphuric acid.}}{SO_2Ho_2} = \underset{\text{Hydric-sodic sulphate.}}{SO_2HoNao} + \underset{\text{Nitric acid.}}{NO_2Ho.}$$

It is very soluble in water, and, like hydric-potassic sulphate, may be used as a flux. When heated it loses water, and forms disodic disulphate, thus:—

$$\underset{\text{Hydric-sodic sulphate.}}{2SO_2HoNao} = OH_2 + \underset{\text{Disodic disulphate.}}{\left\{\begin{array}{l} SO_2Nao \\ O \\ SO_2Nao \end{array}\right.} .$$

The disodic disulphate, when more strongly heated, yields sulphuric anhydride (SO_3) and sodic sulphate:—

$$\underset{\text{Disodic disulphate.}}{\left\{\begin{array}{l} SO_2Nao \\ O \\ SO_2Nao \end{array}\right.} = \underset{\text{Sodic sulphate}}{SO_2Nao_2} + \underset{\text{Sulphuric anhydride.}}{SO_3.}$$

Sodic Carbonate (soda, soda crystals), $CONao_2$.—This most important sodium compound was formerly almost entirely obtained from the ash of burnt seaweeds. The plants whose ash yield it are the *Salsola soda*, grown on the coasts of Spain; *Salicornia annua*, grown on the southern coast of France; and *varec*, a seaweed grown on the rocks on the coast of Brittany. The ash resulting from the combustion of these plants is known as *barilla*, and contains about 18 per cent. of sodic carbonate. The *natron* or *trona* lakes of Egypt have also yielded it from a very remote period. These sources have now, however, almost entirely given place to what is known as *Leblanc's process* of preparing it from sodic chloride (common salt).

By *Leblanc's process* the sodic chloride is first converted into sodic sulphate, as described under that substance; the resulting sulphate is then ground to powder, mixed with ground calcic carbonate and ground coal, in the proportion of 3 parts of sodic sul-

[illegible] carbonate, and 1½ to 2 of coal, and [illegible] charges of 2½ cwt. on to the [illegible] reverberatory furnace, where it is well [illegible] When the mass has melted, effer- [illegible] from the escape of carbonic [illegible] by the carbon of the fuel [illegible] sulphate to sodic sulphide (SNa_2) [illegible] oxygen, thus :—

$$[illegible] = 5SNa_2 + 20CO.$$

Sodic sulphide. Carbonic oxide.

[illegible] as it is evolved burns in jets on [illegible] material with a greenish [illegible]

[illegible] has taken place, a second [illegible] the sodium of the sulphide [illegible] carbonate exchange places, [illegible] calcic sulphide, or pro- [illegible] produced, thus :—

$$[illegible] = 5CaS,2CaO + 2CO_2$$

Calcic* oxysulphide. Carbonic anhydride.

[illegible] the charge is raked out

yield of carbonate, the solution is evaporated to dryness and heated in a furnace with sawdust, which converts the sodic hydrate into carbonate, and the result of the operation is the *soda ash* of commerce, containing about 56 per cent. of available alkali (sodic oxide [ONa_2] in the form of carbonate).

We see, then, that the chemical changes in the manufacture are three in number:—1st, the conversion of sodic chloride into sodic sulphate:—$2NaCl+SO_2Ho_2 = SO_2Nao_2+2HCl$; 2nd, the reduction of the sodic sulphate to sodic sulphide by the carbon of the coal: —$5SO_2Nao_2 + 20C = 5SNa_2 + 20CO$; and 3rd, the decomposition of the sodic sulphide by the calcic carbonate, sodic carbonate and calcic oxysulphide resulting:—$5SNa_2 + 7COCao'' = 5CONao_2 + 5CaS,2CaO + 2CO_2$.

Sodic carbonate has a nauseous, alkaline taste, is very soluble in water, and crystallizes in rhomboidal prisms, often of very large size; the crystals have the composition represented by $CONao_2,10OH_2$, and are efflorescent. It is used in immense quantities in the manufactures of glass, soap, and of the various sodium compounds; large quantities are employed in dyeing and calico printing, scouring wool and yarn, for softening hard waters, and for domestic purposes.

Hydric-sodic Carbonate (bicarbonate of soda), $COHoNao$.—This salt is prepared either by passing carbonic anhydride into an aqueous solution of sodic carbonate, or over moist crystals of that substance; the carbonic anhydride is absorbed, and in the case of the solution the hydric-sodic carbonate precipitates in crystals, which are nearly pure. In the case where the gas is passed over the crystals, heat is evolved, and the crystals lose their water of crystallization. The latter, though the more economical plan, does not yield so pure a product as the former: the change undergone is the same in both cases:—

$$\underset{\substack{\text{Sodic}\\ \text{carbonate.}}}{\mathrm{CONao_2}} + \mathrm{OH_2} + \underset{\substack{\text{Carbonic}\\ \text{anhydride.}}}{\mathrm{CO_2}} = \underset{\substack{\text{Hydric-sodic}\\ \text{carbonate.}}}{2\mathrm{COHoNao}}.$$

Hydric-sodic carbonate is not so soluble in water as the normal carbonate. It crystallizes in rectangular four-sided prisms.

Sodic Nitrate (cubic nitre, Chili saltpetre), $\mathrm{NO_2Nao}$.—This compound occurs in nature in large beds, about a yard below the surface of the soil, in Peru, South America. It is very soluble in water, and is deliquescent.* It is employed in the manufacture of sulphuric and nitric acids, potassic nitrate, and as a manure. The crystals of it are rhombohedral in shape. It cannot be employed as a substitute for potassic nitrate in gunpowder, owing to its deliquescence.

Hydric-disodic Phosphate (rhombic phosphate of sodium, phosphate of soda), $\mathrm{POHoNao_2}$.—This compound is readily prepared by adding a solution of sodic carbonate to a solution of tetrahydric-calcic phosphate, $\mathrm{P_2O_2Ho_4Cao''}$, when calcic carbonate is precipitated, and hydric-disodic phosphate remains in solution:—

$$\underset{\substack{\text{Tetrahydric-calcic}\\ \text{phosphate.}}}{\mathrm{P_2O_2Ho_4Cao''}} + \underset{\substack{\text{Sodic}\\ \text{carbonate.}}}{2\mathrm{CONao_2}} = \underset{\substack{\text{Hydric-disodic}\\ \text{phosphate.}}}{2\mathrm{POHoNao_2}} + \underset{\substack{\text{Calcic}\\ \text{carbonate.}}}{\mathrm{COCao''}} + \underset{\substack{\text{Carbonic}\\ \text{anhydride.}}}{\mathrm{CO_2}} + \underset{\text{Water.}}{\mathrm{OH_2}}.$$

By evaporating the filtered solution, the phosphate may be obtained in rhombic prisms, which are efflorescent. Its aqueous solution has a slightly alkaline reaction. The crystals have the composition $POHoNao_2,12OH_2$. Its solution corrodes flint-glass bottles, causing white flakes of silicic acid to separate. It is used in the laboratory as a test for magnesium.

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* Liable to turn liquid from absorption of moisture from the *atmosphere.*

Sodic Hyposulphite, $SS''ONao_2$.—This salt is prepared by first forming sodic sulphite ($SONao_2$), by passing a stream of washed sulphurous anhydride (SO_2) into a strong solution of sodic carbonate:—

$$\underset{\text{Sodic carbonate.}}{CONao_2} + SO_2 = \underset{\text{Sodic sulphite.}}{SONao_2} + CO_2.$$

The solution of the sulphite is then digested for several days with sulphur, by which means the sulphite takes up sulphur, and is converted into hyposulphite:—

$$\underset{\text{Sodic sulphite.}}{2SONao_2} + S_2 = \underset{\text{Sodic hyposulphite.}}{2SS''ONao_2.}$$

On evaporating the solution, crystals of the hyposulphite, having the composition $SSONao_2,5OH_2$, separate out. It is largely used for photographic purposes as a fixing agent, as it dissolves the silver salts which have been unacted on by the light, and so prevents the portrait or picture from becoming black.

Sodic Borate, $BNao_3$.—This compound does not exist in nature; but an abnormal sodic borate, known as *borax*, and having the composition $B'''_4O_5Nao_2, 10OH_2$, is contained in the mineral called *tincal*, obtained from the East Indies. It is also obtained from certain lakes in Persia and California, and large quantities are prepared by adding sodic carbonate to the native boric acid(BHo_3) of Tuscany:—

$$\underset{\text{Boric acid.}}{4BHo_3} + CONao_2 = \underset{\text{Borax.}}{B_4O_5Nao_2} + 6OH_2 + CO_2.$$

It is soluble in water, and crystallizes in six-sided prisms; its solution has an alkaline reaction on test-paper. When the crystals are heated they fuse in their water of crystallization, bubble up, owing to the escape of steam, and afterwards undergo igneous fusion to a transparent glass. This glass, since it dissolves metallic compounds with production of various characteristic colours, is often employed as a blowpipe test for metals in qualitative analysis. (See p. 86.)

This same property of dissolving metallic compounds which borax possesses is made of use in the welding of metals. The pieces to be joined are heated in the forge, sprinkled with borax, which dissolves the coat of metallic oxide, and when sufficiently heated the pieces are removed, a blow of the hammer or rap against the anvil causes the dissolved oxide to fall off, and the two clean metallic surfaces are speedily united by the blows of the striker. Borax is used in glazing earthenware, as a flux in the smelting and refining of metals, and for softening water.

Sodic Silicate (soluble glass), $SiNao_4$.—This compound is prepared—(1) by fusing together sodic carbonate and silicic anhydride (SiO_2),—

In admixture with other silicates, sodic silicate occurs in glass, and it (equally with potassic silicate) imparts the property of viscidity before fusion to such mixtures, which is of such value in the working of glass.

Sodic Tungstate, WO_2Nao_2.—This compound is employed for rendering fabrics non-inflammable, and to prevent the serious accidents, and in many cases loss of life, which occur every year, through ladies' dresses, &c., accidentally taking fire. A concentrated neutral solution of the salt diluted with about one-third of its bulk of water, and then mixed with three per cent. of hydric-disodic phosphate, is the solution used for this purpose.

Tests for Sodium Compounds.—Sodium compounds are all, with the exception of the antimoniate, very soluble in water; therefore they cannot be precipitated by any reagent, and consequently their presence has to be determined almost entirely by negative evidence. A solution from which all metals, save potassium, sodium, and ammonium, have been precipitated, and which on evaporation to dryness and ignition leaves a residue of saline matter, must contain either potassium or sodium, or both of them. If a little of this solid residue be dissolved in water, and platinic chloride and a few drops of hydrochloric acid added, and the mixture on shaking yields a yellow precipitate, potassium is *present*, and sodium must be tested for in the remainder of the residue by ascertaining whether it colours the Bunsen flame *yellow*, and gives one bright yellow band in the spectroscope; if so, sodium is *present* also. If the platinic chloride gives no precipitate, then potassium is *absent* and sodium is *present*, and its presence may be confirmed by observing the flame coloration, as before stated.

Silver—Ag_2.

Atomic weight, 108. Molecular weight (probable), 216.
Atomicity, I. Sp. gr., 10·47.

Special Notice.—Silver was known to the ancients, being one of the first-discovered metals.

Occurrence in Nature.—Silver occurs in nature in the free state to some extent: in combination it occurs as sulphide (SAg_2), known as *silver glance*, also as chloride ($AgCl$), and bromide ($AgBr$) in *embolite;* it also occurs as sulphide mixed with plumbic sulphide (PbS) in *galena.* The chief sources of its ores are the Hartz mountains in Germany, Kongsberg in Norway, and the Andes in South America.

Preparation.—In Germany the ores are first roasted with sodic chloride, by which they are converted into argentic chloride ($AgCl$), thus:—

$$\underset{\substack{\text{Argentic}\\\text{sulphide.}}}{SAg} + 2NaCl + \underset{\substack{\text{From}\\\text{the air.}}}{2O_2} = \underset{\substack{\text{Argentic}\\\text{chloride.}}}{2AgCl} + \underset{\substack{\text{Sodic}\\\text{sulphate.}}}{SO_2Nao_2}.$$

The chloride is mixed with water, scrap iron, and into sacks, which act as rude filters; the greater portion of the mercury drains off into a trough placed underneath, leaving a soft solid amalgam, containing 15 to 17 per cent. of silver. This is placed on trays, in charges of 5 cwt. at a time, under a large distillatory bell of iron, and a fire is made round the upper part of the latter; the mercury volatilizes, descends in vapour, and is condensed by cold water placed in a trough round the mouth of the bell. A spongy mass of silver, containing 70 per cent. silver, 28 per cent. copper, and other impurities amounting to 2 per cent., is left, and is afterwards melted and cast into ingots. In South America, where fuel is expensive, the metal is extracted by a somewhat different process, the chemistry of the operation being rather obscure.

Silver is often contained in lead ores to such an extent that it is profitable to extract it even when it is not in greater proportion than three or four ounces of silver to a ton of lead. This is done by a process discovered by Mr. Pattinson, of Newcastle; the argentiferous lead is melted in a cast-iron pot, the fire is then withdrawn and the metal stirred briskly; the alloy of silver and lead, being more fusible than the lead alone, remains in a melted state, while the lead crystallizes out, and is removed as it forms by means of an iron ladle. This operation is repeated on the more concentrated argentiferous lead until the lead contains about 300 ounces of silver to the ton, beyond which point it is not found advantageous to carry the concentration. To obtain the silver free from the lead it is now submitted to the operation known as *cupellation*. The lead is melted in a reverberatory furnace, in which a cupel made of bone-ash is fitted, and a blast of air passing over its surface causes that metal to oxidize;

the oxide so formed is absorbed by the cupel, while the silver, together with a small proportion of lead, remains on its surface, five tons being reduced to three cwt.; this is run off while melted, and when a sufficient quantity has accumulated, it is submitted a second time to the operation, and the silver is obtained free from lead.

Silver may be precipitated from any of its solutions, whether acid or alkaline, by means of the metal magnesium. This plan has recently been adopted by many photographers for recovering silver from their residues.

The pure metal may be obtained by dissolving ordinary silver (which contains copper) in nitric acid,—

$$2Ag_2 + \underset{\text{Nitric acid.}}{6NO_2Ho} = \underset{\substack{\text{Argentic}\\ \text{nitrate.}}}{4NO_2Ago} + \underset{\substack{\text{Nitrous}\\ \text{anhydride.}}}{N_2O_3} + 3OH_2;$$

adding sodic chloride to the solution, which precipitates argentic chloride,—

$$2NO_2Ago + \underset{\text{Sodic chloride.}}{2NaCl} = \underset{\substack{\text{Argentic}\\ \text{chloride.}}}{2AgCl} + \underset{\text{Sodic nitrate.}}{2NO_2Nao};$$

washing the precipitate by decantation with water, and then boiling it with solution of sodic hydrate (NaHo) and a few lumps of cane sugar. The boiling is continued until the powder becomes completely soluble in nitric acid; the liquid is then poured off, and the metallic powder washed by decantation: the residue is pure silver, and may be at once used for the preparation of silver salts, or it may be melted and cast into an ingot.

Properties.—Silver is a brilliant white metal, and does not oxidize or lose its brilliancy in pure air. Even when melted in air it does not oxidize, but absorbs mechanically about 22 times its bulk of oxygen, which it gives up again on becoming solid, giving rise to a phenomenon called the "spitting" of the silver.

It is malleable and ductile, and probably the best conductor of heat and electricity. In air containing vapour of sulphur or sulphuretted hydrogen it tarnishes, becoming coated with the sulphide (SAg_2), which is black. This is the reason why silversmiths usually burn their gas outside their shops.

Uses.—Silver is extensively used in making plate, in coining, electro-plating, &c. For the two former uses it is always alloyed with copper, pure silver being too soft to resist wear and tear sufficiently. English coins are made of silver alloyed with 7½ per cent. copper, and French ones contain 10 per cent.

Principal Compounds.—The salts of silver are without colour if the constituent acid be so, and leave in many cases metallic silver after ignition. Silver compounds are decomposed by magnesium, zinc, iron, copper, ferrous sulphate, and stannous chloride, metallic silver being precipitated.

Argentic Oxide, OAg_2—This oxide may be prepared by adding sodic hydrate (NaHo) to a solution of argentic nitrate,—

$$\underset{\text{Sodic hydrate.}}{NaHo} + \underset{\text{Argentic nitrate.}}{NO_2Ago} = \underset{\text{Argentic hydrate.}}{AgHo} + \underset{\text{Sodic nitrate.}}{NO_2Nao};$$

and heating the hydrate thus produced to 60°,—

$$\underset{\text{Argentic hydrate.}}{2AgHo} = \underset{\text{Argentic oxide.}}{OAg_2} + \underset{\text{Water.}}{OH_2}.$$

Argentic Chloride, $AgCl$.—When a soluble chloride is added to a solution of silver salt, this compound is always precipitated, in the form of a white curdy solid. It is quite insoluble in water and nitric acid. It is sometimes found native, and is then called *horn* *silver.* It is decomposed by exposure to light, especially if any organic substances are present. It is reduced to the metallic state when put into water with metallic zinc or iron and a little sulphuric acid. It is soluble in ammonic hydrate, potassic cyanide ($\mathbf{C}NK$), and in the sodic hyposulphite ($\mathbf{S}S''ONao_2$).

On the blackening of argentic chloride and other silver salts, by exposure to light, depends the process of photography as now ordinarily practised. When heated, argentic chloride melts at about 260°, and sublimes at a higher temperature; if a current of pure and dry hydrogen be passed over it while fused, it is reduced to metallic silver with formation of hydrochloric acid:—

$$\underset{\text{Argentic chloride.}}{2AgCl} + \underset{\text{Hydrogen.}}{H_2} = \underset{\text{Silver.}}{Ag_2} + \underset{\text{Hydrochloric acid.}}{2HCl.}$$

It is also decomposed by sodic carbonate when fused with it, metallic silver being produced:—

$$\underset{\text{Argentic chloride.}}{4AgCl} + \underset{\text{Sodic carbonate.}}{2\mathbf{C}ONao_2} = \underset{\text{Silver.}}{2Ag_2} + \underset{\text{Sodic chloride.}}{4NaCl} + \underset{\text{Carbonic anhydride.}}{2\mathbf{C}O_2} + \underset{\text{Oxygen.}}{O_2.}$$

Argentic Iodide, AgI.—When a soluble iodide is added to a solution of a silver salt, this salt is precipitated as a pale yellow flocculent solid:—

$$\underset{\text{Argentic nitrate.}}{\mathbf{N}O_2Ago} + \underset{\text{Potassic iodide.}}{KI} = \underset{\text{Argentic iodide.}}{AgI} + \underset{\text{Potassic nitrate.}}{\mathbf{N}O_2Ko.}$$

It occurs to some extent in nature in Mexico. It is insoluble in nitric acid, and nearly so in ammonic *hydrate* ($\mathbf{N}H_4Ho$); in the latter respect being an exception to silver salts generally. It is an important agent in the photographic art.

Argentic Nitrate (lunar caustic, nitrate of silver), NO_2Ago.—This important salt is prepared by dissolving pure silver in nitric acid. The solution, when concentrated by evaporation, yields the salt in square anhydrous tables. When melted and poured into cylindrical moulds it forms the *lunar*[*] *caustic* of the surgeon. When exposed to light, if impure, the salt blackens, from its reduction to metallic silver, especially if any organic matter be present. It stains the skin black, and is employed as a dye for the hair, and for the manufacture of marking-ink for linen. Preparations for photography and electro-plating are also made from it.

Argentic Sulphate, SO_2Ago_2.—When silver is boiled with strong sulphuric acid, sulphurous anhydride (SO_2) is evolved, and argentic sulphate formed :—

$$\underset{\text{Silver.}}{Ag_2} + \underset{\text{Sulphuric acid.}}{2SO_2Ho_2} = \underset{\text{Argentic sulphate.}}{SO_2Ago_2} + \underset{\text{Sulphurous anhydride.}}{SO_2} + \underset{\text{Water.}}{2OH_2}.$$

It requires ninety times its weight of water for solution, and may be crystallized in small rhombic prisms. It is formed in separating silver from gold.

Tests for Silver Compounds.—Hydrochloric acid and soluble chlorides, when added to solutions of silver salts, precipitate the white curdy argentic chloride ($AgCl$), which is soluble in ammonic hydrate, and insoluble in nitric acid. Heated in the reducing blowpipe flame with sodic carbonate on charcoal, silver compounds are readily reduced to metallic silver.

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[*] From *luna*, the moon, the alchemist's name for silver.

Uses.—Owing to the difficulty hitherto experienced in the preparation of this metal, it has not yet become possible to put it to any use. No doubt, when this is overcome and its properties are well known, it will be found of use in some way or other.

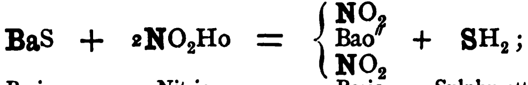
Principal Compounds.

The soluble salts of barium, as well as the carbonate, are very poisonous; the best antidote is magnesic or sodic sulphate.

Baric Oxide (baryta), $\mathbf{Ba}O$.—This substance may be prepared either from the naturally occurring sulphate or carbonate. If the sulphate be employed, it is ground to powder, mixed with ground charcoal, and heated to redness in a crucible, by which means the carbon deprives the sulphate of its oxygen, baric sulphide ($\mathbf{Ba}S$) and carbonic oxide (CO) being formed:—

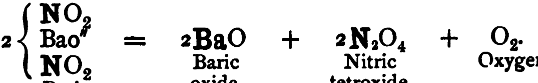
$$\underset{\substack{\text{Baric}\\\text{sulphate.}}}{SO_2Bao''} + \underset{\text{Carbon.}}{4C} = \underset{\substack{\text{Carbonic}\\\text{oxide.}}}{4CO} + \underset{\substack{\text{Baric}\\\text{sulphide.}}}{\mathbf{Ba}S.}$$

The baric sulphide so obtained is decomposed by nitric acid, in order to form the nitrate, thus:—



$$\underset{\substack{\text{Baric}\\\text{sulphide.}}}{\mathbf{Ba}S} + \underset{\substack{\text{Nitric}\\\text{acid.}}}{2NO_2Ho} = \underset{\substack{\text{Baric}\\\text{nitrate.}}}{\left\{\begin{matrix}NO_2\\Bao''\\NO_2\end{matrix}\right.} + \underset{\substack{\text{Sulphuretted}\\\text{hydrogen.}}}{SH_2;}$$

and the nitrate, after evaporation to dryness, is ignited in an iron crucible, by which it is decomposed with evolution of nitric peroxide and oxygen, and formation of baric oxide:—



$$\underset{\substack{\text{Baric}\\\text{nitrate.}}}{2\left\{\begin{matrix}NO_2\\Bao''\\NO_2\end{matrix}\right.} = \underset{\substack{\text{Baric}\\\text{oxide.}}}{2\mathbf{Ba}O} + \underset{\substack{\text{Nitric}\\\text{tetroxide.}}}{2N_2O_4} + \underset{\text{Oxygen.}}{O_2.}$$

When the carbonate is employed, it is first decomposed by nitric acid to form the nitrate:—

$$COBao'' + 2NO_2Ho = \left\{\begin{matrix} NO_2 \\ Bao'' \\ NO_2 \end{matrix}\right. + CO_2 + OH_2;$$

Baric carbonate.	Nitric acid.	Baric nitrate.	Carbonic anhydride.	Water.

and the nitrate is then heated as before stated.

Baric oxide, as thus obtained, is a grey porous mass, which absorbs moisture and carbonic anhydride from the air; when treated with water it combines with it, or slakes with great evolution of heat, forming baric hydrate:—

$$BaO + OH_2 = BaHo_2.$$

Baric oxide.	Water.	Baric hydrate.

Baric oxide is used in the laboratory for absorbing carbonic anhydride in gaseous analyses, and as an alkali where potassic or sodic hydrates are inapplicable.

Baric Peroxide, BaO_2.—If baric oxide or [redacted] by adding a solution of baric hydrate to its solution in very dilute hydrochloric acid; these plates are insoluble in water. When free from water it is a white powder resembling magnesic oxide. If heated to a higher temperature than that at which it is formed, baric peroxide is decomposed, giving up half its oxygen and becoming baric oxide again:—

$$\underset{\substack{\text{Baric}\\\text{peroxide.}}}{2BaO_2} = \underset{\substack{\text{Baric}\\\text{oxide.}}}{2BaO} + O_2.$$

Or if it be treated with steam at the same temperature as it was formed, oxygen is evolved, and baric hydrate is produced:—

$$\underset{\substack{\text{Baric}\\\text{peroxide.}}}{2BaO_2} + \underset{\substack{\text{Water.}\\\text{(as steam).}}}{2OH_2} = \underset{\substack{\text{Baric}\\\text{hydrate.}}}{2BaHo_2} + O_2.$$

When treated with acids, baric peroxide is decomposed, hydroxyl (Ho_2) being formed (see Hydroxyl):—

$$\underset{\substack{\text{Baric}\\\text{peroxide.}}}{BaO_2} + \underset{\substack{\text{Hydrochloric}\\\text{acid.}}}{2HCl} = \underset{\substack{\text{Baric}\\\text{chloride.}}}{BaCl_2} + \underset{\text{Hydroxyl.}}{\left\{\begin{matrix}OH\\OH\end{matrix}\right.}\cdot$$

It has been proposed to prepare oxygen from air by the alternate oxidation of baric oxide and reduction of the baric peroxide so produced; but as yet the process is not practically carried out, owing to certain difficulties which it presents. Baric peroxide is used in the laboratory for preparing hydroxyl, and certain compounds in organic chemistry called the organic peroxides.

Baric Hydrate (hydrate of baryta), $BaHo_2$.—This compound is prepared, as already stated, by the action of water on baric oxide:—$BaO + OH_2 = BaHo_2$; or of barium on water:—$2OH_2 + Ba = BaHo_2 + H_2$;

[illegible] sulphide with cupric [illegible]

[illegible] = [illegible] $+ 4S''Cu'_2$.
[illegible] Cuprous
[illegible] sulphide.

[illegible] cuprous [illegible] baric hydrate, [illegible] separate out. [illegible] water. $BaHo_2$ [illegible] a strong alkaline [illegible] similar to those of [illegible]

[illegible] substance is prepared [illegible] baric [illegible] mixed with an [illegible] weight of [illegible] hour in [illegible] the carbon [illegible] oxygen from the [illegible] oxide being [illegible] lized in lead mines. It has a high sp. gr., viz., 4·6, and the name *baryta*, formerly employed for baric oxide, was given on this account, being derived from the Greek *βαρὺς* (heavy). Baric sulphate is insoluble in water, and nearly so in acids, hence it is produced as a white precipitate when a solution of a sulphate is added to a solution of a barium salt. On account of its not being blackened by sulphuretted hydrogen, it is used by artists in water-colours as a *permanent white.* It is not well fitted for use as an oil-colour, as it becomes partially transparent when ground with oil; it is, however, used to some extent for adulterating white-lead.

Baric Carbonate (witherite, carbonate of baryta), $COBao''$.—This salt, like the sulphate, is abundant in lead mines, in England, Styria, and Siberia. It may be prepared by adding a solution of a carbonate to a solution of a barium salt; it then precipitates as a white powder, insoluble in water:—

$BaCl_2$	$+$	$COKo_2$	$=$	$COBao''$	$+$	$2KCl.$
Baric chloride.		Potassic carbonate.		Baric carbonate.		Potassic chloride.

It is used to some extent in place of potassium, sodium, and lead compounds, in the manufacture of glass, and in the laboratory for preparing barium compounds.

Baric Chloride (muriate of baryta, chloride of barium), $BaCl_2$.—If baric carbonate be decomposed by hydrochloric acid,—

$COBao''$	$+$	$2HCl$	$=$	$BaCl_2$	$+$	CO_2	$+$	$OH_2,$
Baric carbonate.		Hydrochloric acid.		Baric chloride.		Carbonic anhydride.		Water.

and the solution filtered and slowly evaporated, crystals of baric chloride, $BaCl_2,2OH_2$, are deposited. It is, however, usually prepared from baric sulphide, by decomposing it by hydrochloric acid:—

$$BaS + 2HCl = BaCl_2 + SH_2.$$

It crystallizes in flat tables. Its aqueous solution is used as a test for sulphuric acid, whose presence it indicates by forming a white precipitate of baric sulphate, insoluble in nitric acid.

Baric Nitrate (nitrate of baryta), $\left\{\begin{matrix}\mathbf{N}O_2\\ Bao''\\ \mathbf{N}O_2\end{matrix}\right.$.—This salt is prepared in the same way as the chloride, substituting dilute nitric acid for hydrochloric:—

$$(1.)\quad COBao'' + 2\mathbf{N}O_2Ho = \left\{\begin{matrix}\mathbf{N}O_2\\ Bao''\\ \mathbf{N}O_2\end{matrix}\right. + CO_2 + OH_2.$$

Baric carbonate.	Nitric acid.	Baric nitrate.	Carbonic anhydride.	Water.

$$(2.)\quad BaS + 2\mathbf{N}O_2Ho = \left\{\begin{matrix}\mathbf{N}O_2\\ Bao''\\ \mathbf{N}O_2\end{matrix}\right. + SH_2.$$

Baric	Baric	Sulphuretted

Tests for Barium Compounds.—The most certain test for barium compounds is the formation of baric sulphate, insoluble in water, acids, and alkalies, when a soluble sulphate or sulphuric acid is added to their solutions; the precipitate separates *immediately*. If sodic or ammonic citrate is present in a neutral or alkaline solution containing barium, the sulphate is not precipitated on adding a solution of a sulphate; but if hydrochloric acid be then added, the sulphate is precipitated as usual.

Potassic chromate (CrO_2Ko_2) causes a yellow precipitate of baric chromate (CrO_2Bao'') in neutral and alkaline barium solutions; the precipitate is insoluble in potassic, sodic, and ammonic hydrates, and acetic acid.

Ammonic oxalate, $\left\{\begin{matrix}COAmo\\COAmo\end{matrix}\right.$, precipitates only from concentrated solutions the white baric oxalate, $\left\{\begin{matrix}CO\\Bao''\\CO\end{matrix}\right.$, soluble in acids.

Barium compounds tinge the Bunsen flame of a light green colour, and when examined by the spectroscope are distinguished by several green bands occurring in the spectrum.

Strontium—Sr.

Atomic weight, 87·5. Molecular weight (probable), 87·5.
Atomicity, II. Sp. gr., 2·5.

Special Notice.—Strontium was discovered by Sir H. Davy immediately after he discovered barium in 1809. It is very similar to barium in its properties.

Occurrence in Nature.—Strontium is a less abundant element than barium. It occurs, similarly to barium, in lead mines, as sulphate (SO_2Sro''), commonly called *celestine;* and carbonate ($COSro''$), known as *strontianite.* The latter name is derived from Strontian, in Argyllshire, near which town it occurs both massive and in crystals.

Preparation.—Strontium may be prepared by the same processes as are described for preparing barium.

Properties.—Strontium is a malleable, pale yellow metal. When heated in air it burns with a crimson flame, forming strontic oxide:—

$$Sr_2 + O_2 = \underset{\text{Strontic oxide.}}{2SrO.}$$

It decomposes water with evolution of hydrogen, and formation of strontic hydrate:—

$$\underset{\text{Water.}}{2OH_2} + Sr = \underset{\text{Strontic hydrate.}}{SrHo_2} + H_2.$$

Uses.—As it has not been produced in quantity yet, strontium in the metallic state has not been put to any use.

$$\underset{\text{Strontic oxide.}}{SrO} + \underset{\text{Water.}}{OH_2} = \underset{\text{Strontic hydrate.}}{SrHo_2}.$$

It is similar to baric hydrate in its properties and uses.

Strontic Chloride (muriate of strontia), $SrCl_2$.—This salt is prepared like baric chloride. It is a deliquescent salt, very soluble in water, as also in alcohol. It crystallizes in needles as $SrCl_2,6ON_2$.

Strontic Sulphate (sulphate of strontia, celestine), SO_2Sro''.—This salt occurs in nature crystallized in the same form as baric sulphate; the density, however, of it is much lower, being 3·9. It is only slightly soluble in water and acids. It may be obtained as a white precipitate by adding a solution of a sulphate to a solution of a strontium salt; unless the solutions are concentrated the precipitate does not separate immediately.

Strontic Carbonate (carbonate of strontia, strontianite), $COSro''$.—As already stated, this salt occurs in nature. It is very slightly soluble in water. It may be obtained by adding a solution of a carbonate to a solution of a strontium salt. It is used for preparing other strontium compounds.

Strontic Nitrate (nitrate of strontia), $\left\{\begin{matrix}NO_2\\ Sro''\\ NO_2\end{matrix}\right.$.—This salt is prepared by decomposing strontic carbonate or sulphide by dilute nitric acid:—

$$(1.)\ \underset{\text{Strontic carbonate.}}{COSro''} + \underset{\text{Nitric acid.}}{2NO_2Ho} = \underset{\text{Strontic nitrate.}}{\left\{\begin{matrix}NO_2\\ Sro''\\ NO_2\end{matrix}\right.} + \underset{\substack{\text{Carbonic}\\ \text{anhydride.}}}{CO_2} + \underset{\text{Water.}}{OH_2}.$$

$$(2.)\ \underset{\substack{\text{Strontic}\\ \text{sulphide.}}}{SrS} + 2NO_2Ho = \left\{\begin{matrix}NO_2\\ Sro''\\ NO_2\end{matrix}\right. + \underset{\substack{\text{Sulphuretted}\\ \text{hydrogen.}}}{SH_2}.$$

It crystallizes in octahedra which are anhydrous. It is moderately soluble in cold water. It is used for preparing strontic oxide, and for preparing a *red fire* for pyrotechnical purposes.

Tests for Strontium Compounds.—Sulphuric acid and soluble sulphates cause the precipitation of the sulphate after agitation and allowing the mixture to stand for some time: the precipitate is insoluble in acids and alkalies.

Potassic chromate (CrO_2Ko_2) produces no precipitate (difference from barium), strontic chromate (CrO_2Sro) being soluble.

Ammonic oxalate, $\begin{cases} COAmo \\ COAmo \end{cases}$, precipitates white strontic oxalate, even from dilute solutions.

Strontium compounds tinge the Bunsen flame crimson, and when examined by the spectroscope are distinguished by giving several bands in the red and one in the blue part of the spectrum.

Calcium.—Ca.

mountain ranges, either alone or combined with magnesic carbonate, forming *magnesian limestone* in the latter case. *Calc spar*, *marble*, *chalk*, *coral*, &c., are names by which the carbonate is known.

Preparation.—Calcium may be prepared by the same processes as barium, and in addition by the following:—(1) By fusing together calcic chloride ($CaCl_2$), sodium, and zinc; the sodium decomposes the calcic chloride, liberating calcium, which then forms an alloy with the zinc:—

$$\underset{\text{Calcic chloride.}}{CaCl_2} + \underset{\text{Sodium.}}{Na_2} = \underset{\text{Sodic chloride.}}{2NaCl} + \underset{\text{Calcium.}}{Ca.}$$

The alloy is then heated in a gas carbon crucible to a very high temperature, which volatilizes the zinc; the calcium, mixed with a little iron, remaining. (2) By fusing together calcic iodide (CaI_2) and sodium:—

$$\underset{\text{Calcic iodide.}}{CaI_2} + Na_2 = \underset{\text{Sodic iodide.}}{2NaI} + Ca.$$

Properties.—Calcium is a malleable, light, yellow metal. Dry air oxidizes it very speedily when the surface is continually exposed to its action, forming calcic oxide (CaO); it also oxidizes in moist air, forming calcic hydrate ($CaHo_2$). Heated in air it burns with a brilliant white flame, forming calcic oxide. It decomposes water at the ordinary temperature, forming calcic hydrate and evolving hydrogen:—

$$\underset{\text{Water.}}{2OH_2} + Ca = \underset{\text{Calcic hydrate.}}{CaHo_2} + \underset{\text{Hydrogen.}}{H_2.}$$

Strong nitric acid does not act on it unless boiling;

the dilute acid dissolves it readily. When heated and plunged into chlorine, or the vapours of bromine, iodine, and sulphur, it burns, evolving very bright light. Mercury readily dissolves it, forming calcium amalgam.

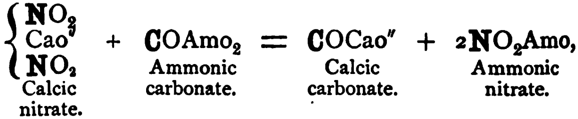
Uses.—Owing to the circumstance that it has not yet been produced in quantity at a moderate price, calcium has not been put to any use.

Principal Compounds.

Calcic Oxide (lime, quicklime, protoxide of calcium), **Ca**O.—This oxide is obtained by heating the carbonate :—

$$\begin{array}{ccccc} \mathbf{C}\text{OCao}'' & = & \mathbf{Ca}\text{O} & + & \mathbf{C}\text{O}_2. \\ \text{Calcic} & & \text{Calcic} & & \text{Carbonic} \\ \text{carbonate.} & & \text{oxide.} & & \text{anhydride.} \end{array}$$

The black limestone of Kilkenny and Derbyshire is the best for this purpose. If lime be required absolutely pure, calcic nitrate is decomposed by ammonic carbonate,—



$$\begin{array}{ccccccc} \left\{\begin{array}{l} \mathbf{N}\text{O}_2 \\ \text{Cao}'' \\ \mathbf{N}\text{O}_2 \end{array}\right. & + & \mathbf{C}\text{OAmo}_2 & = & \mathbf{C}\text{OCao}'' & + & 2\mathbf{N}\text{O}_2\text{Amo}, \\ \text{Calcic} & & \text{Ammonic} & & \text{Calcic} & & \text{Ammonic} \\ \text{nitrate.} & & \text{carbonate.} & & \text{carbonate.} & & \text{nitrate.} \end{array}$$

and the precipitated calcic carbonate ignited to whiteness in a crucible. Lime for ordinary purposes is obtained by calcining chalk or limestone (impure calcic carbonate) in a kiln; a red heat continued for some hours is sufficient to liberate the whole of the carbonic anhydride. In the best limekilns the process is carried on continuously by broken limestone and fuel being thrown in at the top, and the lime raked out at intervals at the bottom. When the limestone contains silica (SiO_2), and the temperature is very great, the lime refuses to slake, calcic silicate ($SiCao''_2$) having been formed; the lime is then said to be *overburned*. Pure lime is a white infusible solid; at a high temperature it emits a pure white light, and is used for this purpose in the oxyhydrogen light. When sprinkled with water it slakes and crumbles into a soft white powder, which is calcic hydrate ($CaHo_2$), the temperature being raised sufficiently to explode gunpowder and char wood: carts laden with lime have sometimes been set on fire by a shower of rain.

Calcic oxide is very largely used in the preparation of mortar and cements. In making mortar the lime is *slaked* with water (*i. e.*, is converted into calcic hydrate), and one part of this paste is mixed with three or four parts of moderately fine sharp sand (silicic anhydride, SiO_2). The reason why mortar hardens is not thoroughly understood; the portions exposed to the air absorb carbonic anhydride from it, and part of the silicic anhydride unites with the lime, forming calcic silicate; the central portions, however, still remain as hydrate, even after thousands of years. Besides this, calcic oxide is largely used in preparing bleaching powder, potassic chlorate and hydrate, sodic hydrate, in purifying gas and water, as a manure, and in making crucibles, which are infusible by the most intense heat known.

Calcic Peroxide, CaO_2.—This compound is prepared, like the corresponding strontium one, by adding hydroxyl to a solution of calcic hydrate:—

$$\underset{\substack{\text{Calcic}\\\text{hydrate.}}}{CaHo_2} + \underset{\text{Hydroxyl.}}{O_2H_2} = \underset{\substack{\text{Calcic}\\\text{peroxide.}}}{CaO_2} + \underset{\text{Water.}}{2OH_2}.$$

Calcic Hydrate (slaked lime, hydrate of lime),

$CaHo_2$.—As already stated, this compound is produced when water is added to calcic oxide; this is the mode of preparing it. This substance is an exception to the rule that substances are more soluble in hot than cold water, for one part of it requires 1,300 parts of boiling water to dissolve it, but only 730 parts of cold. The solution is known as *lime water*, and if exposed to the air absorbs carbonic anhydride from it, forming calcic carbonate, which forms as a white film on the surface of the solution:—

$$\underset{\text{Calcic hydrate.}}{CaHo_2} + \underset{\text{Carbonic anhydride.}}{CO_2} = \underset{\text{Calcic carbonate.}}{COCao''} + \underset{\text{Water.}}{OH_2}.$$

This formation of calcic carbonate is used as a test for the presence of carbonic anhydride in gaseous mixtures. If calcic hydrate be heated to redness, calcic oxide is reproduced:—

$$\underset{\text{Calcic hydrate.}}{CaHO_2} = \underset{\text{Calcic oxide.}}{CaO} + \underset{\text{Water.}}{OH_2}.$$

Calcic hydrate, mixed with water to form a thin paste, is called *milk of lime*.

Calcic Carbonate (chalk, limestone, marble, Iceland spar, aragonite, calc spar, carbonate of lime), $COCao''$.—This is one of the most widely distributed compounds in the mineral world, forming whole mountains of immense thickness in almost every part of the world. It is the chief constituent of chalk, shells, and coral reefs. In nature calcic carbonate is seldom pure; it generally contains clay, ferric oxide, and organic matter. Limestone caverns are often lined with magnificent crystals of calcareous spar, which have been slowly *deposited* from a watery solution. These stalactitic incrustations are often of great beauty. In pure water calcic carbonate is scarcely soluble, but when the water contains free carbonic acid it is taken up on account of the formation of the dihydric-calcic carbonate, which is soluble in water:—

$$\underset{\text{Calcic carbonate.}}{COCao''} + \underset{\text{Carbonic acid.}}{COHo_2} = \underset{\text{Dihydric-calcic carbonate.}}{C_2O_2Ho_2Cao''}.$$

If lime water be poured into a vessel containing carbonic anhydride, a milky white precipitate of calcic carbonate is formed; but on agitation, if the carbonic anhydride be in excess, the turbidity disappears, and a clear, transparent solution of dihydric-calcic carbonate is obtained. If the clear solution is boiled, carbonic anhydride is expelled, and calcic carbonate again precipitates. All natural waters contain carbonic acid. This soluble carbonate is almost always present in river and spring water. In limestone districts it is often present to a very great extent. The *hardness* of water is usually owing to its presence (see page 77). Tea-kettles and boilers in which such water is heated often become lined with a thick incrustation. Dr. Ritterband has proposed a method of preventing the formation of this deposit in boilers, which consists of throwing into the boiler a small quantity of ammonic chloride, when ammonic carbonate is formed, which is volatilized by the high temperature, calcic chloride remaining in solution. This method is objectionable on account of the corrosive action of the ammonic carbonate on the machinery, and is not applicable to permanently hard waters, which generally contain calcic sulphate. All the other salts of calcium may be obtained from the carbonate by decomposing it by the various acids.

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It crystallizes as $CaCl_2,6OH_2$, in six-sided prisms, which produce cold on solution in water. It is a product of many chemical operations, and, owing to its great attraction for water, is used for drying gases; for absorbing the water produced by combustion of hydrogen in organic analyses; it is also used as a reagent.

Calcic Chloro-hypochlorite (bleaching powder, chloride of lime), $Ca(OCl)Cl$.—This substance has already been described (see page 81).

If bleaching powder be heated its bleaching properties are entirely destroyed, and if it be exposed to the air in an open vessel it absorbs carbonic anhydride and evolves chlorine. It thus varies in value with its age, the care bestowed on its preparation, and its subsequent exclusion from the atmosphere.

The general outline of the method pursued in bleaching is as follows:—The goods are soaked in a dilute solution of bleaching powder, they are then removed to a vat containing dilute sulphuric acid. The bleaching powder now undergoes decomposition by the acid, chlorine being liberated:—

$$Ca(OCl)Cl + SO_2HO_2 = SO_2Cao'' + OH_2 + Cl_2.$$

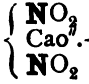
Bleaching powder. Sulphuric acid. Calcic sulphate. Water. Chlorine.

The chlorine in contact with the article to be bleached causes a complete destruction of the colouring matter. It acts by decomposing water, setting free oxygen, which oxidizes the colour:—

$$2OH_2 + 2Cl_2 = O_2 + 4HCl.$$

Water. Chlorine. Oxygen. Hydrochloric acid.

As the solutions are very dilute, this process has to be repeated several times (see properties of chlorine, p. 61).

Calcic Nitrate (nitrate of lime), 

$\left\{\begin{matrix}NO_2\\ Cao''\\ NO_2\end{matrix}\right.$.—This salt is prepared by decomposing calcic carbonate by dilute nitric acid :—

$$COCao'' + 2NO_2Ho = \left\{\begin{matrix}NO_2\\Cao''\\NO_2\end{matrix}\right. + OH_2 + CO_2.$$

Calcic carbonate. Nitric acid. Calcic nitrate. Water. Carbonic anhydride.

It is a deliquescent salt, crystallizing in prisms having the composition $\left\{\begin{matrix}NO_2\\Cao''\\NO_2\end{matrix}\right.4OH_2$. It is soluble in water and alcohol.

Calcic Phosphate (bone ash, phosphate of lime, coprolite, apatite, sombrerite, phosphorite), $P_2O_2Cao''_3$.—This salt is found in nature as *coprolite*, *apatite*, *phosphorite*, and *sombrerite*, and it forms the principal constituent of bones. It may be obtained as a white precipitate by adding a solution of a phosphate to a solution of a calcium salt :—

$$3CaCl_2 + 2PONao_3 = 6NaCl + P_2O_2Cao''_3.$$

Calcic chloride. Sodic phosphate. Sodic chloride. Calcic phosphate.

mous quantities as a bye-product in the manufacture of sodic carbonate, and is known as *soda waste*.

Calcic Fluoride (fluor spar, Derbyshire spar), $\mathbf{Ca}F_2$.—This salt occurs in nature in lead mines, being found in Derbyshire, Cornwall, and Cumberland, both crystallized and massive. It varies in colour according as it is mixed with one or other of various metallic oxides. It is very slightly soluble in water. If a soluble fluoride or hydrofluoric acid is added to a solution of a calcium salt, calcic fluoride precipitates as a white gelatinous powder:—

$$\underset{\text{Calcic chloride.}}{CaCl_2} + \underset{\text{Hydrofluoric acid.}}{2HF} = \underset{\text{Calcic fluoride.}}{CaF_2} + \underset{\text{Hydrochloric acid.}}{2HCl.}$$

It is the chief source of the compounds of fluorine; when heated with sulphuric acid it yields hydrofluoric acid (see page 139). It is also used as a flux in smelting operations.

Tests for Calcium Compounds. — Sulphuric acid and soluble sulphates only cause precipitates of calcic sulphate in concentrated solutions, and even in these cases it re-dissolves in excess of water.

Potassic chromate ($\mathbf{Cr}O_2Ko_2$) causes no precipitate, as calcic chromate ($\mathbf{Cr}O_2Cao''$) is soluble.

Ammonic oxalate $\left(\begin{cases} COAmo \\ COAmo \end{cases}\right)$ causes, even in very dilute alkaline solutions, an immediate white precipitate of calcic oxalate, $\begin{cases} CO \\ Cao'' \\ CO \end{cases}$, soluble in acids.

Examined by the spectroscope, calcium compounds are distinguishable by giving a bright orange line and a slightly less bright broad green one, besides several faint ones.

Magnesium—Mg.

Atomic weight, 24. Molecular weight probably, 24.
Specific gravity, [illegible]

History.—Magnesium was shown to exist by Davy, in 1808, but was first isolated by M. Bussy, in 1829.

Occurrence in Nature.—The compounds of this metal are of frequent occurrence in nature; as carbonate ($CO''Mgo''$) it is found in combination with calcic carbonate, forming *dolomite* or *magnesian limestone*; as silicate ($SiMgo''_2$) it occurs as *meerschaum*, *steatite*, *French chalk* or *soapstone*, *talc*, *asbestos*, *amphibole*, and *amianthus*; as sulphate (SO_2Mgo''), known as *Epsom salts*, it occurs in mineral waters; from those of Epsom it derives its name; and as hydrate ($MgHo_2$) in *brucite*.

Preparation.—Magnesium is prepared by the following processes:—(1) By fusing together anhydrous magnesic chloride ($MgCl_2$) and sodium in a clay crucible:—

boiled with it, hydrogen being liberated and magnesic hydrate formed :—

$$\underset{\text{Water.}}{2OH_2} + \underset{\text{Magnesium.}}{Mg} = \underset{\text{Magnesic hydrate.}}{\mathbf{Mg}Ho_2} + \underset{\text{Hydrogen.}}{H_2}.$$

Heated in air, it burns with a very intense white light; the sole product is magnesic oxide :—

$$\underset{\text{Oxygen.}}{O_2} + \underset{\text{Magnesium.}}{Mg_2} = \underset{\text{Magnesic oxide.}}{2\mathbf{Mg}O}.$$

It can be distilled in an atmosphere of hydrogen, being about as volatile as zinc: advantage is taken of this property in order to obtain it pure. It is soluble in hydrochloric acid and ammonic chloride, but is unacted upon by a mixture of concentrated nitric and sulphuric acids unless boiling. It burns brilliantly when heated in chlorine and the vapours of bromine, iodine, and sulphur.

Uses.—The light evolved by magnesium when burnt being very similar to sunlight, the metal is used to some extent for photographic and lighting purposes. It has been found experimentally by Bunsen and Roscoe that a very fine magnesium wire when burning emits a light equal to that given by seventy-four stearine candles of five to the pound. It has been proposed to replace the zinc usually employed in Marsh's apparatus in arsenical poisoning cases by magnesium on account of its greater freedom from arsenic. It displaces silver from its solutions, and is employed by the photographer for recovering that metal from his waste products.

Principal Compounds.

Magnesic Oxide (magnesia, calcined magnesia), **Mg**O.—This substance is the only known oxide of magnesium. It is prepared by igniting either the artificially prepared carbonate or nitrate; or by burning the metal in air or oxygen :—

$$(1.)\quad \underset{\text{Magnesic carbonate.}}{COMgo''} = \underset{\text{Magnesic oxide.}}{MgO} + \underset{\text{Carbonic anhydride.}}{CO_2}.$$

$$(2.)\ 2\underset{\text{Magnesic nitrate.}}{\left\{\begin{matrix}NO_2\\ Mgo''\\ NO_2\end{matrix}\right.} = \underset{\text{Magnesic oxide.}}{2MgO} + \underset{\text{Nitric peroxide.}}{2N_2O_4} + \underset{\text{Oxygen.}}{O_2}.$$

$$(3.)\quad \underset{\text{Magnesium.}}{Mg_2} + \underset{\text{Oxygen.}}{O_2} = \underset{\text{Magnesic oxide.}}{2MgO}.$$

Magnesic oxide is a light, white, infusible, tasteless, and nearly insoluble powder. The oxide obtained from the nitrate is denser than that from the carbonate. When placed upon red litmus paper and moistened with water it turns it blue, showing its alkaline nature. The formation of the hydrate, which occurs on adding water to it, takes place without any evolution of heat :—

$$\underset{\text{Magnesic oxide.}}{MgO} + \underset{\text{Water.}}{OH_2} = \underset{\text{Magnesic hydrate.}}{MgHo_2}.$$

Its principal use is in medicine.

Magnesic Hydrate (hydrate of magnesia), $MgHo_2$.—This is produced, as just stated, by adding water to the oxide. It occurs native in crystalline scales, forming *brucite*.

Magnesic Carbonate (carbonate of magnesia, magnesite, magnesia alba), $COMgo''$.—This compound occurs in nature as *magnesite*, and combined with calcic carbonate forms *magnesian limestone* or *dolomite*, $C_2O_2Mgo''Cao''$. It is prepared artificially by adding a solution of potassic carbonate to a boiling solution of a magnesium salt, and suspending the precipitate which is produced in water through which carbonic anhydride is passing. The precipitate gradually dissolves, and the compound crystallizes out as $COMgo',3OH_2$, as the solution loses part of its carbonic anhydride; the crystals are efflorescent.

Magnesia alba, as the carbonate sold in the shops is called, is a mixture of magnesic carbonate and magnesic hydrate. It may be obtained very pure by *Pattinson's process*, which consists in heating magnesian limestone till the magnesic carbonate becomes magnesic oxide :—

$$\underset{\substack{\text{Magnesian}\\\text{limestone.}}}{C_2O_2Mgo''Cao''} = \underset{\substack{\text{Calcic}\\\text{carbonate.}}}{COCao''} + \underset{\substack{\text{Magnesic}\\\text{oxide.}}}{MgO''} + \underset{\substack{\text{Carbonic}\\\text{anhydride.}}}{CO_2}.$$

It is then mixed with water in a strong iron vessel, and carbonic anhydride under pressure is passed into it until it ceases to be absorbed; this dissolves the magnesic oxide by converting it into dihydric-magnesic carbonate, $C_2O_2,Mgo''Ho_2$, while the calcic carbonate remains undissolved :—

$$\underset{\substack{\text{Mixed calcic}\\\text{carbonate and}\\\text{magnesic oxide.}}}{COCao'' + MgO} + \underset{\text{Water.}}{OH_2} + \underset{\substack{\text{Carbonic}\\\text{anhydride.}}}{2CO_2} = \underset{\substack{\text{Calcic}\\\text{carbonate.}}}{COCao''} + \underset{\substack{\text{Dihydric-magnesic}\\\text{carbonate.}}}{C_2O_2Mgo''Ho_2}.$$

The solution is then run off and boiled; carbonic anhydride escapes, and the magnesic carbonate is deposited, which is drained and dried at a low temperature :—

$$\underset{\substack{\text{Dihydric-magnesic}\\\text{carbonate.}}}{C_2O_2Mgo''Ho_2} = \underset{\substack{\text{Magnesic}\\\text{carbonate.}}}{COMgo''} + \underset{\text{Water.}}{OH_2} + \underset{\substack{\text{Carbonic}\\\text{anhydride.}}}{CO_2}.$$

Its principal use is in medicine.

Magnesic Sulphate (sulphate of magnesia, Epsom salts), SO_2Mgo''.—This compound occurs in many mineral waters. It is prepared in large quantities by the action of sulphuric acid upon magnesian limestone:—

$$\underset{\text{Magnesian limestone.}}{C_2O_2Mgo''Cao''} + \underset{\text{Sulphuric acid.}}{2SO_2Ho_2} = \underset{\text{Calcic sulphate.}}{SO_2Cao''} + \underset{\text{Magnesic sulphate.}}{SO_2Mgo''} + \underset{\text{Water.}}{2OH_2} + \underset{\text{Carbonic anhydride.}}{2CO_2}.$$

When the action of the sulphuric acid has ceased, the liquor is run into a tank, where it remains for three hours, in order that the calcic sulphate may precipitate. It is then evaporated and allowed to crystallize, which it does in rhombic prisms, as $SOHo_2Mgo'', 6OH_2$.

Magnesic sulphate is soluble in three times its weight of water at the ordinary temperature; the solution has a nauseous bitter taste. It is largely used by warp-sizers for increasing the weight of the warps,—as a spècies of fraud, in fact. It is also used in medicine, and as a reagent in the laboratory. When heated, it loses six molecules of water, and is left as SO_2Mgo'',OH_2.

Magnesic Chloride (muriate of magnesia), $MgCl_2$.—This compound is obtained free from water by dissolving magnesic carbonate in hydrochloric acid:—

$$\underset{\text{Magnesic carbonate.}}{COMgo''} + \underset{\text{Hydrochloric acid.}}{2HCl} = \underset{\text{Magnesic chloride.}}{MgCl_2} + \underset{\text{Water.}}{OH_2} + \underset{\text{Carbonic anhydride.}}{CO_2};$$

and to the resulting chloride solution adding ammonic chloride in sufficient quantity to form magnesic-ammonic chloride ($MgCl_2AmCl$). On now evaporating the solution to dryness and igniting, the ammonic chloride sublimes, and anhydrous magnesic chloride remains:—

$$\underset{\text{Magnesic-ammonic chloride.}}{\mathbf{Mg}Cl_2,AmCl} = \underset{\text{Magnesic chloride.}}{\mathbf{Mg}Cl_2} + \underset{\text{Ammonic chloride.}}{AmCl.}$$

This addition of ammonic chloride is necessary, as otherwise the magnesic chloride undergoes decomposition with the water on evaporating, magnesic hydrate and hydrochloric acid being formed :—

$$\underset{\text{Magnesic chloride.}}{\mathbf{Mg}Cl_2} + \underset{\text{Water.}}{2OH_2} = \underset{\text{Magnesic hydrate.}}{\mathbf{Mg}Ho_2} + \underset{\text{Hydrochloric acid.}}{2HCl.}$$

Magnesic chloride is an extremely deliquescent salt; the anhydrous form evolves heat when dissolved in water; it crystallizes from its aqueous solution in needles having the composition $\mathbf{Mg}Cl_2,6OH_2$. When the anhydrous salt is heated it fuses, and solidifies in confused crystals on cooling. It is used in the manufacture of magnesium.

Magnesic Phosphate, $P_2O_2Mgo''_3$.—This phosphate is produced when sodic phosphate ($\mathbf{P}ONao_3$) is added to a solution of a magnesium salt :—

$$\underset{\text{Trisodic phosphate.}}{2\mathbf{P}ONao_3} + \underset{\text{Magnesic chloride.}}{3\mathbf{Mg}Cl_2} = \underset{\text{Magnesic phosphate.}}{\mathbf{P}_2O_2Mgo''_3} + \underset{\text{Sodic chloride.}}{6NaCl.}$$

The more important phosphates of magnesium are, however, *hydric-magnesic phosphate* ($\mathbf{P}OMgo''Ho$) and *ammonic-magnesic phosphate* ($\mathbf{P}OMgo'Amo$). The former is precipitated by adding hydric-disodic phosphate ($\mathbf{P}ONao_2Ho$) to a solution of a magnesium salt :—

$$\underset{\text{Magnesic sulphate.}}{\mathbf{S}O_2Mgo''} + \underset{\text{Hydric-disodic phosphate.}}{\mathbf{P}ONao_2Ho} = \underset{\text{Hydric-magnesic phosphate.}}{\mathbf{P}OMgo''Ho} + \underset{\text{Sodic sulphate.}}{\mathbf{S}O_2Nao_2.}$$

The latter, by adding ammonic chloride and ammonic formed, on addition of hydric disodic phosphate, to a solution of magnesium salt:—

$$SO_4Mg'' + AmCl + PO_4Na_2H + AmHO =$$
Magnesic sulphate. Ammonic chloride. Hydric disodic phosphate. Ammonic hydrate.

$$PO_4Mg''Am + SO_4Na_2 + AmCl + OH_2$$
Ammonio-magnesic phosphate. Sodic sulphate. Ammonic chloride. Water.

Ammonio-magnesic phosphate is often produced in the determination of magnesium in quantitative analysis, as when ignited it parts with ammonia gas and water, leaving magnesic pyrophosphate, $P_2O_7Mg''_2$, which admits of being easily weighed:—

$$2PO_4Mg''Am = 2NH_3 + OH_2 + P_2O_7Mg''_2$$
Ammonio-magnesic phosphate. Ammonia gas. Water. Magnesic pyrophosphate.

Magnesic Silicate. $SiMg''_2$.—Various silicates of Magnesium occur in nature; as instances may be mentioned *meerschaum*, *soapstone*, *talc*, and *asbestos*. Meerschaum is used for making tobacco-pipes; and asbestos, which is fibrous, is used occasionally in hydrate, gives a white precipitate of ammonic-magnesic phosphate ($POMgo''Amo$), which is soluble in acids.

Heated on charcoal, with a little cobaltous nitrate, $\left\{\begin{matrix}NO_2\\Coo''\\NO\end{matrix}\right.$, magnesium compounds acquire a pink colour.

Zinc—Zn.

Atomic weight, 65. Molecular weight, 65. Molecular volume, normal. Atomicity, II. Sp. gr., 6·8 to 7·2.

Special Notice.—Zinc was described by Paracelsus in the sixteenth century. It is often called *spelter* in commerce.

Occurrence in Nature.—Zinc is found in moderate abundance as carbonate ($COZno''$), known as *calamine;* as sulphide (ZnS), known as *zinc blende* or *black jack;* and as zincic oxide (ZnO), called from its red colour (due to ferric and manganic oxides) *red oxide of zinc.* The ores occur in Cumberland, Derbyshire, Flintshire, and Somersetshire, in England; Silesia and Carinthia, in Germany; and New Jersey, in the United States.

Preparation.—In the extraction of zinc, the ore, whether sulphide or carbonate, is first roasted, by which it is converted into oxide:—

$$(1.)\quad \underset{\text{Zincic carbonate.}}{COZno''} = \underset{\text{Zincic oxide.}}{ZnO} + \underset{\text{Carbonic anhydride.}}{CO_2}.$$

$$(2.)\quad \underset{\substack{\text{Zincic}\\\text{sulphide.}}}{2ZnS} + \underset{\substack{\text{Oxygen}\\\text{(from the air).}}}{3O_2} = \underset{\text{Zincic oxide.}}{2ZnO} + \underset{\substack{\text{Sulphurous}\\\text{anhydride.}}}{2SO_2}.$$

The roasting of the sulphide is a tedious process, owing to the length of time required to convert it into oxide. The roasted ore is then mixed with ground coal or coke, and introduced (in England) into crucibles placed in a furnace; through the bottom of the furnace and crucible an iron pipe passes: the carbon of the coal deprives the zincic oxide of its oxygen, forming carbonic oxide, which passes off as gas, and the zinc being converted into vapour, distils, passing down the iron pipe into a vessel containing water, where it condenses :—

$$\underset{\text{Zincic oxide.}}{ZnO} + \underset{\text{Carbon.}}{C} = \underset{\text{Zinc.}}{Zn} + \underset{\text{Carbonic oxide.}}{CO}.$$

In Silesia and Belgium the distillation is carried on in earthenware retorts.

Properties.—Zinc is a bluish white metal, brittle at ordinary temperatures, and exhibiting a crystalline structure when broken. When heated to between 100° and 150° C. it is malleable, and can be rolled into sheets; but if it be heated to 170° to 200° C. it becomes so brittle as to be easily powdered in a mortar. At 500° it melts, and, according to Deville, its boiling point is 1040° C. If exposed to the air near its boiling point it takes fire, burning with a brilliant white light, and forming zincic oxide. It retains its lustre in dry air, but in moist air it is speedily coated with a thin film of oxide, which protects the underlying metal from further oxidation. When boiled with water, zinc slowly decomposes it, evolving hydrogen :—

$$\underset{\text{Water.}}{2OH_2} + \underset{\text{Zinc.}}{Zn} = \underset{\text{Zincic hydrate.}}{ZnHo_2} + \underset{\text{Hydrogen.}}{H_2}.$$

Acids are decomposed by zinc with evolution of hydrogen at the ordinary temperature; and when boiled with the hydrates of potassium, sodium, and ammonium, hydrogen is also evolved, and a mixed oxide produced (see preparation of hydrogen, method 5, page 54).

Uses.—Zinc is largely used, owing to its cheapness, durability, and lightness, as a substitute for lead in roofing. Iron is coated with zinc in order to prevent its oxidation; the zinc is melted and preserved from oxidation by a layer of ammonic chloride on its surface, and the iron (plate or wire) rendered very clean, is dipped into the zinc, which adheres to its surface; the resulting coated plate or wire is known as *galvanized iron*, the name being given from the fact that the galvanic action set up causes the zinc to oxidize before the iron, and so prevents the iron rusting. Zinc enters into the composition of *brass*, *German silver*, and other alloys. In the galvanic battery the positive pole is almost invariably made of zinc. Zinc also forms the basis of one or two paints. In the laboratory it is in frequent use as a reducing agent, and for preparing hydrogen.

Principal Compounds.

Zincic Oxide (oxide of zinc, zinc white), ZnO.—This compound is prepared, as already stated, by burning zinc in air; the zinc is distilled into a chamber through which air is drawn, the metal burns, and the oxide deposits in condensing chambers. It cannot be used as a substitute for *white-lead* in oil paints, as it does not combine with the oil. It is occasionally obtained crystallized in six-sided prisms. When hot it is yellow, but becomes white on cooling. It is substituted with advantage for *red-lead* in making optical glasses. It dissolves in acids, forming zinc salts. It occurs in nature to some extent, forming *red zinc*, or *red oxide of zinc*.

Zincic Hydrate (hydrated oxide of zinc), $ZnHo_2$.—This hydrate is obtained by adding potassic hydrate solution to a solution of zinc salt:—

$$\underset{\text{Zincic sulphate.}}{SO_2Zno''} + \underset{\text{Potassic hydrate.}}{2KHo} = \underset{\text{Potassic sulphate.}}{SO_2Ko_2} + \underset{\text{Zincic hydrate.}}{ZnHo_2}.$$

The hydrate is thus obtained as a white precipitate, soluble in excess of potassic hydrate. Sodic or ammonic hydrate may be substituted for the potassic hydrate.

Zincic Sulphide (zinc blende, sulphide of zinc, black jack), **ZnS**.—This substance is the principal natural compound of zinc. It occurs both massive and crystallized. It is produced artificially as a white gelatinous hydrate, by addition of sulphide of potassium, sodium, or ammonium to a solution of a salt of zinc:—

$$\underset{\text{Zincic sulphate.}}{SO_2Zno''} + \underset{\substack{\text{Ammonic}\\\text{sulphide.}}}{SAm_2} = \underset{\text{Ammonic sulphate.}}{SO_2Amo_2} + \underset{\substack{\text{Zincic}\\\text{sulphide.}}}{ZnS.}$$

It may also be produced with almost explosive violence by rapidly heating a mixture of zinc and mercuric sulphide:—

$$\underset{\text{Mercuric sulphide.}}{HgS} + Zn = \underset{\text{Zincic sulphide.}}{ZnS} + \underset{\text{Mercury.}}{Hg.}$$

Zincic sulphide does not fuse when heated; heated in the air, the sulphur burns off as sulphurous anhydride, and the metal oxidizes to zincic oxide (see preparation of zinc).

The anhydrous sulphide is very slightly acted on by sulphuric and hydrochloric acids, but nitric acid and *aqua regia* speedily decompose it. The hydrated sulphide is readily soluble in acids.

Zincic Chloride (butter of zinc), **ZnCl**$_2$.—When zinc is heated in chlorine or dissolved in hydrochloric acid, the product is zincic chloride,—alone in the first case, and with simultaneous production of hydrogen in the second:—

$$(1.)\quad \underset{\text{Zinc.}}{Zn} + \underset{\text{Chlorine.}}{Cl_2} = \underset{\substack{\text{Zincic}\\\text{chloride.}}}{ZnCl_2.}$$

$$(2.)\quad Zn + \underset{\substack{\text{Hydrochloric}\\\text{acid.}}}{2HCl} = \underset{\substack{\text{Zincic}\\\text{chloride.}}}{ZnCl_2} + \underset{\text{Hydrogen.}}{H_2.}$$

By the first process it is at once obtained in the solid form. The solution obtained in the second case loses water till it acquires the temperature of 250° C., when it becomes anhydrous; it still remains fluid, however, and if it be allowed to cool, it solidifies to a white, highly deliquescent mass, which sublimes at a red heat, but may be heated to 370° C. without giving off any large amount of fumes. It is fusible at 100° C. It is employed in organic research for depriving compounds of the elements of water; as an antiseptic, forming *Burnett's disinfecting fluid;* as a caustic in surgery, it acts very powerfully on the skin; and in a state of fusion, as a hot bath for maintaining a high but measurable temperature. It is also used along with ammonic chloride to dissolve the film of oxide in soldering.

Zincic Carbonate (calamine), $COZno''$.—Zincic carbonate occurs in nature as *calamine*, both in the crystalline state and massive. The word *calamine* is derived from the Latin *calamus*, a reed, and is given to this compound from the fact that after fusion it adheres to the base of the furnace in a reed-like form. It is generally greyish or yellowish in colour, and readily loses carbonic anhydride when heated, becoming zincic oxide:—

$$\underset{\text{carbonate}}{\underset{\text{Zincic}}{COZno''}} = \underset{\text{oxide.}}{\underset{\text{Zincic}}{ZnO}} + \underset{\text{anhydride.}}{\underset{\text{Carbonic}}{CO_2.}}$$

If a solution of sodic carbonate be added to a solution of a zinc salt, zincic carbonate is not the only product, but the precipitate has a variable composition, consisting of a mixture of hydrate and carbonate. It is used in preparing other compounds of zinc.

Zincic Sulphate (white vitriol, sulphate of zinc), SO_2Zno''; crystallized, $SOHo_2Zno'',6OH_2$.—This salt is the residue left in preparing hydrogen by means of zinc and sulphuric acid. It may also be obtained by roasting zincic sulphide at a low temperature, treating the roasted mass with water and evaporating the solution :—

$$\underset{\substack{\text{Zincic}\\\text{sulphide.}}}{\mathbf{Zn}S} \quad + \quad \underset{\substack{\text{Oxygen}\\\text{(from air).}}}{2O_2} \quad = \quad \underset{\substack{\text{Zincic}\\\text{sulphate.}}}{\mathbf{S}O\,Zno''}.$$

Like magnesic sulphate it crystallizes in colourless four-sided prisms; it is consequently said to be *isomorphous* (Greek, ισος, like; μορφη, form) with it. The crystals are efflorescent in dry air, and when heated to a moderate temperature give up six molecules of water; the remaining molecule is driven off at a higher temperature. It is used in calico-printing and as an emetic in medicine.

Tests for Zinc Compounds.—Ammonic sulphide gives a white precipitate of the hydrated sulphide ($\mathbf{Zn}S$). Sulphuretted hydrogen causes no precipitate from acid solutions of zinc salts, except the acetate. The hydrates of potassium, sodium, and ammonium precipitate the hydrate ($\mathbf{Zn}HO_2$), soluble in excess. The carbonates of potassium, sodium, and ammonium precipitate a basic carbonate of variable composition, which is soluble in excess of ammonic carbonate. Heated on charcoal before the blowpipe, zinc compounds are reduced to metallic zinc, which volatilizes, burning and forming white fumes of the oxide. If moistened with cobaltous nitrate and heated on charcoal, a green infusible mass results.

Cadmium—Cd.

Atomic weight, 112. Molecular weight, 112. Molecular volume, normal. 1 litre of cadmium vapour weighs 56 criths. Atomicity, II. Sp. gr., 8·7.

Special Notice.—Cadmium was discovered in 1818, by Stromeyer, in some zincic oxide which had been prepared for medicinal use. He at first concluded it was arsenic, from its yielding a yellow sulphide with sulphuretted hydrogen, but on further examination he found it to be a new element. The name *cadmium* is derived from the Greek *καδμεια* (kadmeia), a term applied to *calamine* and the zincic oxide produced in brass furnaces.

Occurrence in Nature.—Cadmium, as already stated, occurs in zinc ores. It is also found as sulphide (CdS), forming *greenockite.*

Preparation.—Owing to the fact that it is more volatile than zinc, cadmium is obtained in the same operation as that which affords zinc. The cadmium sublimes first and is collected separately.

Properties.—Cadmium is similar to zinc in appearance, and also resembles tin to some extent. It is so soft as to leave marks on paper. It is both ductile and malleable at the ordinary temperature, but when heated to 80° C. it is so brittle as to be easily powdered. It melts at 315° C., and if exposed to the air it burns, producing brownish yellow fumes of cadmic oxide. It is soluble in hot hydrochloric and sulphuric acids with evolution of hydrogen. Nitric acid dissolves it still more freely.

Uses.—Cadmium is of use in the preparation of alloys of low fusibility. An alloy of 3 parts cadmium, 15 of bismuth, 4 of tin, and 8 of lead, forms a silver-white solid of sp. gr. 9·4, which softens between 55° and 60°, and is completely liquid at 60°; it expands slightly on solidifying. Another alloy, consisting of 1 part cadmium, 7 of bismuth, and 6 of lead, melts at 83°.

Principal Compounds.

Cadmic Oxide, CdO.—This oxide is obtained

prepared by digesting together pulverized metallic cadmium and iodine with water, and evaporating the solution :—

$$\underset{\text{Cadmium.}}{Cd} + \underset{\text{Iodine.}}{I_2} = \underset{\text{Cadmic iodide.}}{\mathbf{CdI_2}}.$$

It is easily crystallized ; the crystals have a pearly lustre, and fuse when heated. It is used for iodizing collodion for photographic purposes.

Tests for Cadmium Compounds.—Sulphuretted hydrogen and ammonic sulphide precipitate from solutions of cadmium salts the yellow sulphide ($\mathbf{Cd}S$), which is insoluble in potassic and sodic sulphides, ammonic hydrate, and hydrochloric acid. It is not volatile when heated, thus being distinguishable from arsenious sulphide ($\mathbf{As}_2S_3$), which it otherwise resembles.

Exposed to the reducing action of carbon, by heating in the inner blowpipe flame with sodic carbonate on charcoal, cadmium compounds give a yellowish brown incrustation of cadmic oxide, from the reduction and subsequent oxidation of the metal.

Mercury (Hydrargyrum)—Hg.

Atomic weight, 200. Molecular weight, 200. Molecular volume, normal. 1 litre of mercury vapour weighs 100 criths. Atomicity, II. Sp. gr., 13·59.

Special Notice.—Mercury was known to the ancients. Its scientific name of *hydrargyrum* is derived from the Greek ὕδωρ (*hudor*), water, and ἄργυρον (*arguron*), silver, *liquid-* or *quick*-silver.

Occurrence in Nature.—Mercury occurs native to some extent; the principal ore of it, however, is the sulphide ($\mathbf{Hg}S$), known as *cinnabar*. It is occasionally found as an amalgam with silver, and rarely as mercurous chloride ($\mathbf{Hg}_2Cl_2$), called *calomel*, and mercuric iodide (HgI_2). Mercurial ores are found in clay slate, red sandstone, and the coal measures. The principal localities where they are found are Idria in Transylvania, Germany; Almaden in Spain; California in North America; and Huancavelica in Peru, South America.

Preparation.—At Almaden the metal is obtained from the sulphide by heating it in a furnace well supplied with air; the sulphur burns off as sulphurous anhydride, and the metal passes off in vapour, and is condensed by passing into cold earthenware tubes, called *aludels*:—

$$\underset{\substack{\text{Mercuric}\\\text{sulphide.}}}{HgS} + \underset{\substack{\text{Oxygen}\\\text{(from air).}}}{O_2} = \underset{\substack{\text{Mercury}\\\text{sulphide.}}}{Hg} + \underset{\substack{\text{Sulphurous}\\\text{anhydride.}}}{SO_2.}$$

By this process great loss of mercury is sustained, owing to imperfect condensation.

In Germany the sulphide is distilled with calcic hydrate in cast-iron retorts; the calcium deprives the sulphide of its sulphur; mercury, and calcic sulphide and sulphate resulting:—

$$\underset{\substack{\text{Mercuric}\\\text{sulphide.}}}{2HgS} + \underset{\substack{\text{Calcic}\\\text{hydrate.}}}{4CaHo_2} = \underset{\text{Mercury.}}{4Hg} + \underset{\substack{\text{Calcic}\\\text{sulphide.}}}{3CaS} + \underset{\substack{\text{Calcic}\\\text{sulphate.}}}{SO_2Cao''} + \underset{\text{Water.}}{4OH_2.}$$

The mercury is condensed in receivers containing water. Instead of calcic hydrate, iron borings may be substituted; ferrous sulphide and mercury are then produced:—

$$\underset{\substack{\text{Mercuric}\\\text{sulphide.}}}{HgS} + \underset{\text{Iron.}}{Fe} = \underset{\substack{\text{Ferrous}\\\text{sulphide.}}}{FeS} + \underset{\text{Mercury.}}{Hg.}$$

Properties.—Mercury is a bright, silvery-looking liquid, being the only metal liquid at the ordinary temperature. It freezes in octahedra at — 40° C., and boils at 350° C.; it volatilizes slightly at the ordinary temperature. It does not oxidize in air unless heated to above 300°, when it slowly becomes converted into mercuric oxide:—

$$\underset{\text{Mercury.}}{2\mathrm{Hg}} + \underset{\text{Oxygen.}}{\mathrm{O}_2} = \underset{\text{Mercuric oxide.}}{2\mathrm{HgO}}.$$

It combines directly with chlorine, bromine, iodine, and sulphur. Hydrochloric acid is without action on it; sulphuric acid, when boiled with it, yields sulphurous anhydride and mercuric sulphate; nitric acid attacks it readily, evolving nitric oxide, and forming, according as the metal or acid is in excess, mercurous or mercuric nitrate.

Uses.—The principal uses of mercury are in amalgamation in the extraction of gold and silver; in silvering and gilding as amalgams; in the preparation of *vermilion*, the artificial sulphide; in the construction of barometers, thermometers, and other philosophical instruments; to replace water in the mercurial pneumatic trough; and for preparing mercurial compounds used in medicine and manufactures.

Principal Compounds.

1. *Mercurous Compounds.*

Mercurous Oxide (suboxide of mercury), $'\mathrm{Hg}'_2\mathrm{O}$.—This oxide is obtained as a black powder by rubbing up powdered mercurous chloride with potassic or sodic hydrate in a mortar:—

$$\underset{\substack{\text{Mercurous}\\\text{chloride.}}}{'\mathrm{Hg}'_2\mathrm{Cl}_2} + \underset{\substack{\text{Potassic}\\\text{hydrate.}}}{2\mathrm{KHo}} + \underset{\substack{\text{Mercurous}\\\text{oxide.}}}{'\mathrm{Hg}'_2\mathrm{O}} + \underset{\substack{\text{Potassic}\\\text{chloride.}}}{2\mathrm{KCl}} + \underset{\text{Water.}}{\mathrm{OH}_2}.$$

If exposed to light or gently heated, it decomposes into mercuric oxide and mercury.

Mercurous Chloride (calomel, subchloride of [illegible]). This salt is prepared (1) by sub[illegible] [illegible] parts mercury and 17 of mer[illegible]

$$[illegible] = Hg'_2Cl_2;$$

Mercurous chloride.

[illegible] mercuric sulphate, mer[illegible]

$$[illegible] = Hg'_2Cl_2 + SO_2Nao_2;$$

Mercurous chloride. Sodic sulphate.

[illegible] sodic chloride to a solution [illegible]

$$[illegible] = Hg'_2Cl_2 + 2NO_2Nao.$$

Mercurous chloride. Sodic nitrate.

[illegible] and [illegible] mercurous [illegible] mercuric chloride: it has, [illegible] to free it from that

Mercurous Nitrate (subnitrate of mercury), $\left\{\begin{matrix}NO_2\\ Hg_2o''\\ NO_2\end{matrix}\right.$.—This salt is prepared by acting on an excess of mercury with nitric acid; nitric oxide is evolved at the same time:—

$$8NO_2Ho + 6Hg = 3\left\{\begin{matrix}NO\\ Hg_2o''\\ NO_2\end{matrix}\right. + 4OH_2 + 2N_2O_2.$$

Nitric acid. Mercury. Mercurous nitrate. Water. Nitric oxide.

It is decomposed by water into an insoluble basic and a soluble acid salt. There are several acid and basic mercurous nitrates. It is used in preparing mercurous chloride to some extent, in testing oils, and as a reagent.

2 *Mercuric Compounds.*

Mercuric Oxide (red or nitric oxide of mercury, red precipitate, peroxide of mercury, HgO.—This oxide is obtained (1.) by heating mercury in air to 300°:

$$2Hg + O_2 = 2HgO;$$

Mercury. Oxygen. Mercuric oxide.

(2.) by heating mercuric nitrate:—

$$2\left\{\begin{matrix}NO_2\\ Hgo''\\ NO_2\end{matrix}\right. = 2HgO + 2N_2O_4 + O_2;$$

Mercuric nitrate. Nitric peroxide. Oxygen.

(3.) by decomposing mercuric chloride or nitrate by potassic or sodic hydrate:—

$$HgCl_2 + 2KHo = HgO + 2KCl + OH_2.$$

Mercuric chloride. Potassic hydrate. Mercuric oxide. Potassic chloride. Water.

As thus prepared it is of a red or yellow colour; that having a yellow colour brings about chemical changes more readily than the red; it appears to be an *allotropic* form of the oxide. When heated it becomes black, but regains its usual colour on cooling; if ignited it is decomposed, metallic mercury and oxygen being evolved:—

$$\underset{\text{Mercuric oxide.}}{2\text{HgO}} \quad = \quad \underset{\text{Mercury.}}{2\text{Hg}} \quad + \quad \underset{\text{Oxygen.}}{\text{O}_2}.$$

It is occasionally used as an oxidizing agent, and for preparing small quantities of oxygen. It was from this oxide that Priestley first obtained oxygen.

Mercuric Chloride (corrosive sublimate, bichloride or perchloride of mercury), HgCl_2.—This important salt may be prepared as follows:—

(1.) By the direct union of its elements:—

$$\underset{\text{Mercury.}}{\text{Hg}} \quad + \quad \underset{\text{Chlorine.}}{\text{Cl}_2} \quad = \quad \underset{\text{Mercuric chloride.}}{\text{HgCl}_2}.$$

(2.) By subliming a mixture of 2½ parts of mercuric sulphate and 1 part of sodic chloride:—

$$\underset{\text{Mercuric sulphate.}}{\text{SO}_2\text{Hgo}''} \quad + \quad \underset{\text{Sodic chloride.}}{2\text{NaCl}} \quad = \quad \text{HgCl}_2 \quad + \quad \underset{\text{Sodic sulphate.}}{\text{SO}_2\text{Nao}_2}.$$

Mercuric chloride crystallizes either in needles or octahedra. It melts at 265°, and boils at 295°; its vapour is very acrid and poisonous. In the solid state its flavour is nauseously metallic and bitter. The best antidote for it is to swallow immediately the whites of several raw eggs, with the albumen of which mercuric chloride forms an insoluble compound, and is thus rendered inert. Ammonic hydrate, when added to a solution of it, causes the precipitation of a substance called, from its white colour, *white precipitate*:—

$$\begin{array}{ccccccccc} HgCl_2 & + & 2AmHo & = & NH_2Hg''Cl & + & AmCl & + & 2OH_2. \\ \text{Mercuric chloride.} & & \text{Ammonic hydrate.} & & \text{Mercurammonic chloride. (white precipitate).} & & \text{Ammonic chloride.} & & \text{Water.} \end{array}$$

Corrosive sublimate is a powerful antiseptic, and is consequently used in preserved anatomical preparations, natural history objects, wood, ropes, &c. A solution of 1 part of the salt in 80 parts water is strong enough for most purposes. It is also used in medicine.

Mercuric Nitrate (pernitrate of mercury), $\left\{\begin{matrix} NO_2 \\ Hgo'' \\ NO_2 \end{matrix}\right.$.—This salt is prepared by acting with excess of nitric acid on mercury, or by dissolving mercuric oxide in the acid:—

$$\begin{array}{lccccccccc} (1.) & 3Hg & + & 8NO_2Ho & = & 3\left\{\begin{matrix} NO_2 \\ Hgo'' \\ NO \end{matrix}\right. & + & 4OH_2 & + & N_2O_2. \\ & \text{Mercury.} & & \text{Nitric acid.} & & \text{Mercuric nitrate.} & & \text{Water.} & & \text{Nitric oxide.} \end{array}$$

$$\begin{array}{lccccccc} (2.) & HgO & + & 2NO_2Ho & = & \left\{\begin{matrix} NO_2 \\ Hgo'' \\ NO_2 \end{matrix}\right. & + & OH_2. \\ & \text{Mercuric oxide.} & & & & & & \end{array}$$

Several basic and acid mercuric nitrates are known. If mercuric nitrate is digested with excess of mercury, it is converted into mercurous nitrate:—

$$\begin{array}{ccccc} \left\{\begin{matrix} NO_2 \\ Hgo'' \\ NO_2 \end{matrix}\right. & + & Hg & = & \left\{\begin{matrix} NO_2. \\ Hg_2o'' \\ NO_2 \end{matrix}\right. \\ \text{Mercuric nitrate.} & & \text{Mercury.} & & \text{Mercurous nitrate.} \end{array}$$

Mercuric nitrate is used for preparing mercuric oxide.

Mercuric Sulphide (protosulphide of mercury, cinnabar, vermilion), **HgS**.—This substance is the most abundant ore of mercury. It is of a scarlet colour. It is prepared either by a *dry* or *wet* process; the *dry* process consists in throwing an intimate mixture of 1 part sulphur and 6 parts mercury, in small portions at a time, into red-hot earthen pots, which are then closed for 32 hours; the pots are then broken, and the sulphide is ground to powder. The *wet* process consists in digesting together mercury, sulphur, and solution of potassic pentasulphide for several days. When sulphuretted hydrogen is passed in excess into a solution of a mercuric salt, mercuric sulphide is obtained as a black precipitate, which becomes scarlet when sublimed:—

$$\underset{\text{Mercuric chloride.}}{HgCl_2} + \underset{\text{Sulphuretted hydrogen.}}{SH_2} = \underset{\text{Mercuric sulphide.}}{HgS} + \underset{\text{Hydrochloric acid.}}{2HCl.}$$

It usually occurs in nature in fibrous masses, but occasionally it is found in hexahedral prisms. It is largely used as a pigment, being called *vermilion.*

Mercuric Sulphate (persulphate of mercury), **SO_2Hgo''**.—This salt is obtained by boiling together two parts of mercury and three of sulphuric acid: mercurous sulphate is first formed, and afterwards decomposed into mercuric sulphate, sulphurous anhydride being evolved the whole time:—

$$(1.)\ \underset{\text{Mercury.}}{2Hg} + \underset{\text{Sulphuric acid.}}{2SO_2Ho_2} = \underset{\text{Mercurous sulphate.}}{SO_2Hg_2o''} + \underset{\text{Water.}}{2OH_2} + \underset{\text{Sulphurous anhydride.}}{SO_2.}$$

$$(2.)\ \underset{\text{Mercurous sulphate.}}{SO_2Hg_2o''} + 2SO_2Ho_2 = \underset{\text{Mercuric sulphate.}}{2SO_2Hgo''} + 2OH_2 + SO_2.$$

It is obtained as a white crystalline powder, which is decomposed by water into a yellow insoluble basic alt, called *turpeth mineral* ($SHgo''_3$) and a soluble acid ne.

Mercuric sulphate is used in the preparation of nercurous and mercuric chlorides.

Mercuric Iodide (red, or biniodide of mercury), HgI_2.—This salt is obtained as a yellow precipitate, which becomes scarlet by exposure to light, by adding potassic iodide to a solution of mercuric chloride :—

$$\underset{\text{Mercuric chloride.}}{HgCl_2} + \underset{\text{Potassic iodide.}}{2KI} = \underset{\text{Mercuric iodide.}}{HgI_2} + \underset{\text{Potassic chloride.}}{2KCl.}$$

The quantity of the two salts used must be calculated according to the proportions given in the above equation, as the iodide is soluble in an excess of either salt. It is used in medicine, and in the preparation of *Nessler's test* for ammonia, which is an aqueous solution of potassic iodide saturated with mercuric iodide, and rendered powerfully alkaline with potassic or sodic hydrate. It produces a yellow or brown coloration, with a trace of ammonia. "In the whole range of chemical analysis there is no determination which surpasses that of ammonia in point of delicacy. It is questionable indeed whether any other approaches it. The Nessler test is capable of indicating one part of ammonia in 20,000,000 parts of water" (Wanklyn and Chapman).

Tests for Mercurous Compounds.—Potassic hydrate precipitates the black oxide ($'Hg'_2O$). Hydrochloric acid and soluble chlorides precipitate the white mercurous chloride, which is blackened by solutions of potassic, sodic, and ammonic hydrates in excess.

Tests for Mercuric Compounds.—Potassic hydrate precipitates the red oxide (HgO). Sulphuretted hydrogen gives first a dirty white precipitate, which changes [illegible] the black sulphide (HgS); [illegible] the sulphide at once; it is [illegible] hydrochloric acid. Hydro- [illegible] gives no precipitate, [illegible]

[illegible] mercuric compounds are re- [illegible] following reactions:—If introduced [illegible] carbonate and heated, [illegible] deposited on the cool part of [illegible] which a piece of gold [illegible] mercurial solution. A white [illegible] amalgamated with [illegible] this stain will [illegible] stannous chloride ($SnCl_2$) [illegible] the metallic state, and [illegible] mercurous chloride and then to [illegible]

$$[illegible] = Hg_2Cl_2 + SnCl_4$$

Mercurous chloride. Stannic chloride.

$$[illegible] = Hg_2 + SnCl_4$$

It derives its scientific name of *cuprum* from the isle of Cyprus, where the Greeks mined and smelted it.

Occurrence in Nature.—Copper often occurs in nature in the metallic state, the largest masses found being over 150 tons in weight; they were obtained from the mines on the shore of Lake Superior, United States of America. The most abundant ore, however, is *copper pyrites*, which is a double sulphide of copper and iron ($Fe_2Cu_2S_4$); but besides it, copper occurs as cuprous oxide, forming *red* or *ruby copper ore* $\left(\left\{\begin{matrix}Cu\\Cu\end{matrix}\right.O\right)$; as cuprous sulphide (*copper glance*, $'Cu'_2S$); cupric sulphide (CuS, *indigo, or blue copper*); *malachite*, or hydrated cupric carbonate ($CO[OCuHo]_2$), &c. The localities for copper ores are Cornwall, Siberia, Hungary, Fahlun, and the district of Lake Superior, United States.

Preparation.—The principal place where copper-smelting is carried on in Great Britain is Swansea, in South Wales. The operations for extracting the metal from the *pyrites* are very complex; but the object in view is to convert the ore into moderately pure cuprous sulphide ($'Cu'_2S$), called *fine metal*, by causing the iron to form a fusible silicate or slag with silica, which is added to the ore when roasted; the sulphide is then roasted to convert it partially into cupric oxide (CuO),—

$$\underset{\text{Cuprous sulphide.}}{'Cu'_2S} + \underset{\text{Oxygen (from air).}}{2O_2} = \underset{\text{Cupric oxide.}}{2CuO} + \underset{\text{Sulphurous anhydride.}}{SO_2};$$

and the mixed sulphide and oxide are then heated, without access of air, which causes the two to react on each other, copper and sulphurous anhydride being formed:—

$$\begin{array}{ccccccc} {}'Cu'_2S & + & 2CuO & = & 4Cu & + & SO_2. \\ \text{Cuprous sulphide.} & & \text{Cupric oxide.} & & \text{Copper.} & & \text{Sulphurous anhydride.} \end{array}$$

Properties.—Copper is a red metal, with a disagreeable odour and taste when rubbed after moistening. It is more malleable than ductile, and is very tenacious. When heated nearly to melting it becomes brittle, and may be easily powdered. It is slightly volatile at a high temperature. Its melting point, estimated by Daniel's pyrometer, is 1090° C. It is one of the best known conductors of heat and electricity. It does not decompose water, even when heated to redness. Heated in the air it oxidizes, forming cupric oxide:—

$$\begin{array}{ccccc} 2Cu & + & O_2 & = & 2CuO. \\ & & & & \text{Cupric oxide.} \end{array}$$

If finely divided it burns like tinder. If copper-foil be placed in chlorine it takes fire spontaneously, forming cupric chloride:—

$$\begin{array}{ccccc} Cu & + & Cl_2 & = & CuCl_2. \\ & & & & \text{Cupric chloride.} \end{array}$$

when hot, cupric sulphate and sulphurous anhydride are produced :—

$$\mathrm{Cu + 2SO_2Ho_2 = SO_2Cuo'' + 2OH_2 + SO_2.}$$

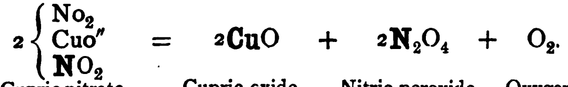
Sulphuric acid. Cupric sulphate. Water. Sulphurous anhydride.

Introduced into the flame of a Bunsen's burner, copper tinges it green.

Uses.—Copper is largely used in sheathing ships, in making boilers for marine engines, vats for dyers, &c. The alloy* now used for the bronze coinage, consists of 95 parts copper, 4 parts tin, and 1 part zinc; *brass* consists of about 64 parts copper and 36 zinc; *gun-metal*, of 90 parts copper and 10 of tin; and *bronze*, of 91 parts copper, 6 parts of zinc, 2 of tin, and 1 of lead.

Cupric Compounds.

Cupric Oxide (black oxide of copper, copper scales, copper oxide, copper monoxide, protoxide of copper), **Cu**O.—This oxide is prepared by the following processes :—(1) By igniting cupric nitrate :—



$$2\left\{\begin{matrix}\mathrm{No_2}\\ \mathrm{Cuo''}\\ \mathrm{NO_2}\end{matrix}\right. = \mathrm{2CuO} + \mathrm{2N_2O_4} + \mathrm{O_2.}$$

Cupric nitrate. Cupric oxide. Nitric peroxide. Oxygen.

(2) By igniting cupric carbonate :—

$$\mathrm{COCuo'' = CuO + CO_2.}$$

Cupric carbonate. Carbonic anhydride.

(3) By decomposing cupric sulphate by potassic hydrate, and boiling the blue hydrate, which is precipitated, in water :—

---

* An *alloy* is a mixture of metals.

$$(a)\ \underset{\text{Cupric sulphate.}}{SO_2Cuo''} + \underset{\text{Potassic hydrate.}}{2KHo} = \underset{\text{Cupric hydrate.}}{CuHo_2} + \underset{\text{Potassic sulphate.}}{SO_2Ko_2}.$$

$$(b)\ (CuHo_2 = Cu + \underset{\text{Water.}}{OH_2}.$$

Cupric oxide so obtained is a black powder, which quickly absorbs moisture from the air. It is soluble in oils, and hence may be the cause of accidental poisoning, if the copper and brass pans used in culinary operations are not kept free from it. When heated it does not give up its oxygen; but if hydrogen or coal gas be passed over it while hot, it gives up its oxygen to them, with evolution of light and heat:—

$$\underset{\text{Cupric oxide.}}{CuO} + \underset{\text{Hydrogen.}}{H_2} = \underset{\text{Copper.}}{Cu} + \underset{\text{Water.}}{OH_2}.$$

When fused with glass it colours it of a beautiful green colour. It dissolves in most acids, forming cupric salts. It is used in organic analysis to supply oxygen

$$(2.)\begin{cases} a.\ \underset{\substack{\text{Cupric}\\ \text{oxide.}}}{CuO} + \underset{\substack{\text{Hydrochloric}\\ \text{acid.}}}{2HCl} = \underset{\substack{\text{Cupric}\\ \text{chloride.}}}{CuCl_2} + \underset{\text{Water.}}{OH_2}. \\ b.\ \underset{\substack{\text{Cupric}\\ \text{carbonate.}}}{COCuo''} + 2HCl = CuCl_2 + OH_2 + \underset{\substack{\text{Carbonic}\\ \text{anhydride.}}}{CO_2}. \end{cases}$$

It crystallizes in green deliquescent needles, having the composition $CuCl_2,2OH_2$. Its solution, if concentrated, is green; when dilute it is blue; and the solid salt, when deprived of its water of crystallization, is somewhat coppery in its hue. It fuses when heated, and at a red heat is converted into cuprous chloride, losing half its chlorine :—

$$\underset{\text{Cupric chloride.}}{2CuCl_2} = \underset{\text{Cuprous chloride.}}{'Cu'_2Cl_2} + \underset{\text{Chlorine.}}{Cl_2}.$$

It is very soluble in alcohol; the solution burns with an intense green flame.

Cupric Sulphate (sulphate of copper, bluestone, blue vitriol), SO_2Cuo''.—This, the most important salt of copper, is prepared (1) by boiling together copper and sulphuric acid (see p. 243 for equation); (2) by roasting or natural oxidation of copper pyrites; (3) by dissolving cupric oxide in sulphuric acid,—

$$\underset{\text{Cupric oxide.}}{CuO} + \underset{\text{Sulphuric acid.}}{SO_2Ho_2} = \underset{\text{Cupric sulphate.}}{SO_2Cuo''} + \underset{\text{Water.}}{OH_2};$$

and (4) by decomposing argentic sulphate (in refining silver) solution by plates of copper,—

$$\underset{\text{Argentic sulphate.}}{SO_2Ago_2} + \underset{\text{Copper.}}{Cu} = \underset{\text{Cupric sulphate.}}{SO_2Cuo''} + \underset{\text{Silver.}}{Ag_2}.$$

Cupric sulphate crystallizes as $SOHo_3Cuo'',4OH_2$ in large blue rhombs. At the boiling point of water it loses four molecules of water, and at 204° C. it is [illegible] a white powder, which becomes [illegible] with water. In taking up this water [illegible] evolves heat. It is largely used [illegible] the formation of *Scheele's* and [illegible] cupric arsenites, *which see*), and [illegible] basic cupric sulphate occurs in [illegible] known as *brochantite*. $SHo_2(OCu''Ho)_2$.

[illegible] [illegible] copper), $\left\{\begin{matrix}NO_2\\ Cuo''\\ NO_2\end{matrix}\right.$.—This [illegible] dissolving copper in nitric acid [illegible] it forms blue deliques- [illegible] strongly it is converted into [illegible]. It is very soluble in [illegible]

[illegible], [illegible].—This normal salt [illegible]. A basic salt known as *ma-* [illegible] dihydrate, $CO[OCu''Ho]_2$) [illegible] beautifully marked green [illegible] high polish, and is [illegible] ornamental purposes. If printing wall-papers, and for dyeing; but it is a dangerous substance to use for such purposes, as it easily rubs off the paper and cloth, and will occasion arsenical poisoning if inhaled with the breath. It may easily be detected by introducing bits of the paper or cloth into Marsh's apparatus (see p. 163). *Schweinfurth green* is a mixture of cupric acetate and cupric arsenite.

Cuprous Compounds.

Cuprous Oxide (suboxide of copper, cupreous oxide, red oxide of copper), $'Cu'_2O$.—This oxide occurs in nature, forming *red* or *ruby copper ore*. It is prepared artificially (1) by boiling a solution of cupric sulphate, to which an excess of potassic hydrate has been added, with sugar; the sugar acts as a reducing agent, and the cuprous oxide precipitates as a red powder; (2) by heating together cupric oxide and copper:—

$$\underset{\text{Cupric oxide.}}{Cu''O} + Cu = \underset{\text{Cuprous oxide.}}{'Cu'_2O}.$$

Its principal use is to stain glass of a ruby-red colour.

Cuprous Hydride (hydride of copper), Cu_2H_2.—This compound is prepared by adding hypophosphorous acid (POH_2Ho) to a solution of cupric sulphate; the hydride precipitates as a brown powder. When heated to 70° C. it decomposes, hydrogen and copper being produced:—

$$\underset{\text{Cuprous hydride.}}{Cu_2H_2} = Cu + H_2.$$

If introduced into chlorine it takes fire spontaneously:—

$$\underset{\substack{\text{Cuprous}\\\text{hydrate.}}}{Cu_2H_2} + \underset{\text{Chlorine.}}{2Cl_2} = \underset{\substack{\text{Cuprous}\\\text{chloride.}}}{Cu_2Cl_2} + \underset{\substack{\text{Hydrochloric}\\\text{acid.}}}{2HCl}.$$

Acted on by hydrochloric acid, it is converted into cuprous chloride with evolution of hydrogen :—

$$\mathbf{Cu}_2\text{H}_2 + 2\text{HCl} = \mathbf{'Cu'}_2\text{Cl}_2 + 2\text{H}_2.$$

This evolution of hydrogen proves that this gas, when in the free or molecular state, is not in the form of a single atom, but of two atoms, the bond of the one satisfying the bond of the other.

Cuprous Sulphide (cupreous sulphide, subsulphide of copper), $\mathbf{'Cu'}_2\text{S}$.—This sulphide forms the mineral called copper glance, which occurs both massive and in six-sided prisms. It may be prepared by heating together 3 parts sulphur and 8 parts copper; the combination takes place with incandescence :—

$$2\text{Cu}_2 + \text{S}_2 = \underset{\text{Cuprous sulphide.}}{2\mathbf{'Cu'}_2\text{S}.}$$

It forms the *fine metal* produced in copper-smelting. It fuses easily, and is soluble in nitric acid and *aqua regia.*

Cuprous Chloride (cupreous chloride, subchloride of copper), $\mathbf{'Cu'}_2\text{Cl}_2$.—This salt is prepared (1) by digesting together metallic copper and cupric chloride,—

$$\underset{\text{Cupric chloride.}}{\mathbf{Cu}\text{Cl}_2} + \text{Cu} = \underset{\text{Cuprous chloride.}}{\mathbf{'Cu'}_2\text{Cl}_2;}$$

(2) by dissolving a mixture of cupric oxide and copper in hydrochloric acid,—

$$\underset{\text{Cupric oxide.}}{\mathbf{Cu}\text{O}} + \text{Cu} + \underset{\text{Hydrochloric acid.}}{2\text{HCl}} = \underset{\text{Cuprous chloride.}}{\mathbf{'Cu'}_2\text{Cl}_2} + \underset{\text{Water.}}{\mathbf{O}\text{H}_2;}$$

(3) by distilling copper with twice its weight of mercuric chloride,—

$$2\text{Cu} + \underset{\text{Mercuric chloride.}}{\mathbf{Hg}\text{Cl}_2} = \mathbf{'Cu'}_2\text{Cl}_2 + \text{Hg}.$$

Cuprous chloride is a white solid, insoluble in water, luble slightly in strong hydrochloric acid; the chlo- le is precipitated on diluting the acid with water. ıe solution of it in hydrochloric acid absorbs carbonic ide gas (CO), forming a compound which crystallizes pearly scales. If the hydrochloric acid solution is posed to the air, oxygen is absorbed, and a pale bluish een precipitate, known as *Brunswick green*, used as a int, precipitates. The compound occurs in nature *atacamite.*

Tests for Copper Compounds.—*Cuprous Salts ıd Compounds.*—Heated on charcoal in inner blow- pe flame with sodic carbonate they yield a bead of etallic copper. They are insoluble in water, and luble in hydrochloric acid: this solution absorbs rbonic oxide. Having been little studied, the sts for them are not thoroughly known.

Cupric Salts and Compounds.—The colour of their lutions is blue or green. Sulphuretted hydrogen ıd ammonic sulphide precipitate the black hydrated lphide (CuS), almost insoluble in ammonic sul- ıide. Potassic and sodic hydrates in excess preci- tate cupric hydrate ($CuHo_2$), which when boiled th the liquid becomes black from the formation of pric oxide (CuO). Ammonic hydrate causes the me precipitate; but when in excess it redissolves it, rming a splendid dark blue solution, very charac- ristic of copper. Potassic ferrocyanide gives a dark ırple-red precipitate of cupric ferrocyanide, insoluble hydrochloric acid. A piece of bright iron dipped to a slightly acidulated solution of a cupric salt comes coated with red metallic copper. Heated th sodic carbonate on charcoal in inner blowpipe me a bead of copper is obtained.

TRIAD METALS.

Gold—Au_2.

Atomic weight, 196·7. Atomicity, III. and I. Sp. gr., 19·3 to 19·5.

Occurrence in Nature.—This metal is widely found in very small quantities throughout nature. It is always obtained in the metallic state, beautifully crystallized in a cubic form, associated with sand, quartz, and oxide of iron. The sands of various rivers contain it in a finely divided state. This is the *gold-dust* of commerce.

Preparation.—The quartz containing the gold is ground to powder, agitated with mercury, which takes up the gold, forming *gold amalgam*, and the amalgam is afterwards distilled, as in the process for preparing silver.

Properties.—Gold is a rich yellow metal, possessing great lustre. It is about as hard as lead, is very ductile, and the most malleable of the metals. The thinnest gold leaf is not more than 1-280,000th of an inch thick. It melts at 1100° C. It does not oxidize in either dry or moist air at any temperature, and is unacted on by any single acid except selenic (SeO_2Ho_2); the solvent for it is a mixture of nitric and hydrochloric acids, which on this account is called *aqua regia* (royal water). Any mixture evolving or containing free chlorine will dissolve it. When potassic or sodic hydrates are fused on it the metal is not corroded, and hence it is used for making crucibles in which these substances can be fused. At the ordinary temperature it combines directly with chlorine and bromine, and with phosphorus when heated.

Uses.—Gold is used for gilding on porcelain and wood, and in the galvanic battery for electro-gilding on metals. Bohemian ruby glass is coloured by gold. It is also largely used for coining. Gold coins are alloyed with a proportion of silver or copper to increase their hardness and durability. English standard gold contains 1-12th of copper. Gold forms a semi-solid amalgam with mercury.

Auric Compounds.

Auric Chloride (terchloride of gold, gold terchloride), $\mathbf{Au}Cl_3$.—When gold is dissolved in *aqua regia* this salt is produced :—

$$Au_2 + \overbrace{\underset{\text{Nitric acid.}}{2\mathbf{N}O_2Ho} + \underset{\text{Hydrochloric acid.}}{6HCl}}^{\text{Aqua regia.}} = \underset{\text{Auric chloride.}}{2\mathbf{Au}Cl_3} + \underset{\text{Water.}}{4\mathbf{O}H_2} + \underset{\text{Nitric oxide.}}{\left\{\begin{matrix}\mathbf{N}O\\ \mathbf{N}O\end{matrix}\right.}$$

This is the most important compound of gold, and is a deliquescent yellow solid, soluble in water, alcohol, and ether. Ether removes it from its aqueous solution. It forms a crystalline compound with hydrochloric acid ($\mathbf{Au}Cl_3,HCl$). It is used in gilding, both *electrotype* and *dipping*. For *dipping* the chloride is mixed with hydric-sodic carbonate in excess, and the articles, after being cleansed with dilute nitric acid, are boiled in the solution. It is also extensively used in photography for giving the tone to paper pictures.

Auric Oxide, $\mathbf{Au}_2O_3$, is also known.

Aurous Chloride (protochloride of gold, gold monochloride), AuCl.—Auric chloride when heated to 185° C. fuses, and loses a molecule of chlorine, forming aurous chloride :—

$$\underset{\text{Auric chloride.}}{\mathbf{Au}Cl_3} = \underset{\text{Aurous chloride.}}{AuCl} + \underset{\text{Chlorine.}}{Cl_2}.$$

It is a yellow, unstable, slightly soluble powder.

Boiled with water it is decomposed into auric chloride and gold:—

$$\underset{\text{Aurous chloride.}}{3AuCl} = \underset{\text{Gold.}}{Au_2} + \underset{\text{Auric chloride.}}{AuCl_3}.$$

Heated to 204° C. it is decomposed into its elements.

Aurous Oxide, $''Au'_2O$, is also known.

Tests for Gold Compounds.—Gold compounds are recognised by the following tests:—On addition of ferrous sulphate (SO_2Feo'') to acidulated solutions free from nitric acid, a brown precipitate of metallic gold is produced. A bar of metallic tin causes a precipitate of the *purple of Cassius*. Oxalic acid $\left(\begin{cases}COHo\\COHo\end{cases}\right)$ also reduces gold compounds to the metallic state.

TETRAD METALS.

Aluminium—Al.

Atomic weight, 27·5. Molecular weight, unknown. Molecular volume, unknown. Atomicity, IV., but always acts as a pseudo-triad. Sp. gr., 2·6.

Special Notice.—Aluminium was discovered by Wöhler in 1828. He obtained it by decomposing aluminic chloride by means of potassium. It derives its name from *alum*, of which it is a constituent.

Occurrence in Nature.—Aluminium is a very abundant element. It occurs in nature in various silicates; *e. g.*, *clay*, *shale*, *kaolin*, *halloysite*, *fuller's-earth*, *albite*, *felspar*, *labradorite*, *mica*, *hornblende*, *gneiss*, *porphyry*, *granite*, *syenite*, *basalt*, *chlorite*, &c.; as hydrate, forming *gibbsite* $\left(\begin{cases}AlHo_3\\AlHo_3\end{cases}\right)$ and *diaspore* $\left(\begin{cases}AlOHo\\AlOHo\end{cases}\right)$; as oxide, forming *corundum*, *sapphire*, and *ruby* $\left(\left\{\begin{matrix}\mathbf{Al}O\\ \mathbf{Al}O\end{matrix}\right.O\right)$; as magnesic aluminate, forming *spinelle* $\left(\left\{\begin{matrix}\mathbf{Al}\\ \mathbf{Al}\end{matrix}\right.O_2Mgo''\right)$; as basic sulphate, forming *aluminite*, $SO_2(Al_2O_2Ho_4)'',7OH_2$, and *alunite* or *alum-stone;* and as fluoride, combined with sodic fluoride, forming *cryolite* ($\mathbf{Al}_2F_6,6NaF$).

Preparation.—Aluminium may be obtained (1) by introducing into a platinum tube, closed at one end, aluminic chloride ($\mathbf{Al}_2Cl_6$), with an equal quantity of potassium loosely wrapped up in platinum-foil. The closed end of the tube is then heated, so as to sublime the chloride, bringing its vapours into contact with the melted potassium:—

$$\underset{\text{Aluminic chloride.}}{\left\{\begin{matrix}'\mathbf{Al}'''Cl_3\\ '\mathbf{Al}'''Cl_3\end{matrix}\right.} + \underset{\text{Potassium.}}{3K_2} = \underset{\text{Aluminium.}}{2Al} + \underset{\text{Potassic chloride.}}{6KCl.}$$

The metal is separated from the potassic chloride by cold water.

(2.) It is now largely prepared from the mineral *cryolite.* Alternate layers of cryolite and sodium are introduced into a crucible and strongly heated; aluminium is separated and sodic fluoride formed:—

$$\underset{\text{Cryolite.}}{'\mathbf{Al}'''_2F_6,6NaF} + \underset{\text{Sodium.}}{3Na_2} = \underset{\text{Sodic fluoride.}}{12NaF} + \underset{\text{Aluminium.}}{2Al.}$$

Properties.—Aluminium is a very light, malleable metal, resembling zinc. It is also ductile, and very sonorous. Heated in air or oxygen to redness it burns, forming aluminic oxide:—

$$4Al + 3O_2 = \underset{\text{Aluminic oxide.}}{2'\mathbf{Al}'''_2O_3.}$$

It melts at a lower temperature than silver, and crystallizes on cooling in regular octahedra. Hydro-

$$Al_2Ho_6 \quad = \quad \left\{\begin{matrix}AlO\\AlO\end{matrix}\right.O \quad + \quad 3OH_2.$$

Aluminic oxide. Water.

Thus obtained, aluminic oxide is a white, tasteless mass, scarcely affected by acids. It is fusible before the oxyhydrogen blowpipe.

Aluminic Hydrate (hydrate of alumina), $\left\{\begin{matrix}AlHo_3\\AlHo_3\end{matrix}\right.$. —*Diaspore* and *gibbsite* are naturally occurring hydrates, the former having the composition $\left(\left\{\begin{matrix}AlOHo\\AlOHo\end{matrix}\right.\right)$, the latter $\left(\left\{\begin{matrix}AlHo_3\\AlHo_3\end{matrix}\right.\right)$.—It is artificially obtained by adding ammonic hydrate to a solution of an aluminium salt (see preceding paragraph for equation). It is a white, bulky, gelatinous powder, soluble in potassic hydrate and acids. In combination with colouring matters it forms the pigments called *lakes*, and it is extensively used as a *mordant* for fixing colours on cloth in dyeing. A modification of aluminic hydrate is known, which is largely soluble in water containing a small quantity of acetic acid.

Sodic Aluminate (aluminate of soda), $\left\{\begin{matrix}AlNao_3\\AlNao_3\end{matrix}\right.$. —This compound is prepared by heating *bauxite*, a mineral containing 75 per cent. of aluminic oxide. This is heated with sodic carbonate until a portion of the mass no longer effervesces with an acid; it is then treated with water, filtered, and evaporated to dryness. It forms a white, infusible solid, freely soluble in water. It is used instead of alum in preparing lakes and in dyeing. It is converted into aluminic hydrate and sodic carbonate by passing carbonic anhydride through it.

When solutions of sodic aluminate and aluminic chloride are mixed, they undergo mutual decomposition, aluminic hydrate and sodic chloride being formed, and the former precipitating:—

$$\underset{\text{Sodic aluminate.}}{\left\{\begin{matrix}AlNao_3\\AlNao_3\end{matrix}\right.} + \underset{\text{Aluminic chloride.}}{Al_2Cl_6} + \underset{\text{Water.}}{6OH_2} = \underset{\text{Aluminic hydrate.}}{2\left\{\begin{matrix}AlHo_3\\AlHo_3\end{matrix}\right.} + \underset{\text{Sodic chloride.}}{6NaCl}$$

Aluminic Chloride (chloride of aluminium), Al'''_2Cl_6 or $\left\{\begin{matrix}AlCl_3\\AlCl_3\end{matrix}\right.$.—This compound cannot be obtained free from water by dissolving aluminic oxide in hydrochloric acid and evaporating the solution to dryness, as the acid is expelled during the evaporation. It is therefore prepared as follows:—Aluminic oxide is mixed with charcoal and strongly calcined in a covered crucible. The contents of the crucible are then transferred to a porcelain tube heated to redness in a furnace; a current of chlorine is passed through the tube, when the aluminic chloride condenses in the cold part of the apparatus:—

$$\underset{\text{Aluminic oxide.}}{\left\{\begin{matrix}AlO\\AlO\end{matrix}\right.O} + \underset{\text{Chlorine.}}{3Cl_2} + \underset{\text{Carbon.}}{3C} = \underset{\text{Aluminic chloride.}}{\left\{\begin{matrix}AlCl_3\\AlCl_3\end{matrix}\right.} + \underset{\text{Carbonic oxide.}}{3CO.}$$

Aluminic chloride is a deliquescent salt, and hisses when thrown on water. The solution on evaporation yields crystals having the composition $Al_2Cl_6,12OH_2$. It is used in the manufacture of aluminium.

Aluminic Fluoride, $\left\{\begin{matrix}AlF_3\\AlF_3\end{matrix}\right.$ or Al'''_2F_6.—This salt occurs, combined with sodic fluoride, in nature, and is obtained in this form in large quantities from Greenland; it is known as *cryolite*. It is used in the manufacture of aluminium, of sodic aluminate, and *hot-cast porcelain.*

Aluminic Sulphate (sesquisulphate of aluminium, aluminium sulphate, concentrated alum), $S_3O_6(Al_2O_6)^{vi},18OH_2$.—This salt is obtained by boiling sulphuric acid with burnt clay, or by saturating dilute sulphuric acid with aluminic hydrate and evaporating, when thin crystallized scales of aluminic sulphate are obtained; it has an acid reaction, and when heated to redness is decomposed. It combines with the sulphates of sodium, potassium, and ammonium, forming double salts. A basic aluminic sulphate, known as *aluminite,* is found in nature; it is aluminic sulphate-tetrahydrate ($SO_2[Al_2O_2Ho_4]'',7OH_2$).

Dipotassic-aluminic Tetrasulphate (alum, potash alum, aluminium potash alum, sulphate of aluminium and potassium), $S_4O_8Ko_2(Al_2O_6)^{vi},24OH_2$.—Alum is manufactured (1) from a slaty kind of clay which contains ferric disulphide (FeS_2). This *alum schist,* as it is called, is first gently roasted, then moistened with water and exposed to the atmosphere. The sulphur of the disulphide is converted into sulphuric acid by combination with oxygen from the air. Ferrous sulphate and aluminic sulphate are produced. These salts are dissolved by lixiviation with water, and the solution is concentrated by evaporation, and mixed with potassic chloride, which decomposes the iron salt, forming ferrous chloride and potassic sulphate. This potassic sulphate then combines with the aluminic sulphate to form alum. The ferrous chloride, which is very soluble in water, is separated from the alum by crystallization :—

$$a.\ \underset{\substack{\text{Ferric}\\\text{disulphide.}}}{2FeS_2} + \underset{\text{Oxygen.}}{7O_2} + \underset{\text{Water.}}{2OH_2} = \underset{\substack{\text{Ferrous}\\\text{sulphate.}}}{2SO_2Feo''} + \underset{\substack{\text{Sulphuric}\\\text{acid.}}}{2SO_2Ho_2}.$$

b. $\mathrm{Si_3O_8(Al_2O_6)''} + 3\mathrm{SO_2Ho_2} = \mathrm{S_3O_6(Al_2O_6)''} + 3\mathrm{SiOHo_2}$
Aluminic silicate. Sulphuric acid. Aluminic sulphate. Silicic acid.

c. $\mathrm{S_3O_6(Al_2O_6)''} + \mathrm{SO_2Ko_2} + 24\mathrm{OH_2} =$
Aluminic sulphate. Potassic sulphate. Water.

$\mathrm{S_4O_8Ko_2(Al_2O_6)''},24\mathrm{OH_2}$
Alum.

(2) By adding potassic sulphate to aluminic sulphate, produced as described under that substance.

Alum crystallizes in regular octahedra, which melt in their water of crystallization when heated; and if the heat be continued, a white, infusible, porous mass of anhydrous or *burnt alum* (the *alumen ustum* of the druggist) is left. Alum is largely used in dyeing as a mordant: the aluminic hydrate, which easily separates when cloth or fibre is put into a solution of it, attaches itself both to the colouring matter and the fibre, and so fixes the colour on the cloth. It is also used in purifying water.

The term alum has now become generic, as it is found that chromium, manganese, and iron may each octahedron; they cannot therefore be separated by crystallization, for they crystallize on each other's crystals.

Ammonia alum is now largely manufactured, and is gradually superseding common potash alum.

Aluminic Silicates.—As already stated, these silicates are numerous; they are also important, as, owing to their abundance and peculiar properties, they are of great use for many purposes. All *clays* are aluminic silicates, but their composition is very variable. The following analysis of the best *Stourbridge fireclay* will show to some extent their composition:—

Silica (SiO_2)	72·04
Aluminic oxide (Al_2O_3) . .	18·79
Water (OH_2)	6·70
Ferric oxide (Fe_2O_3) . . .	1·68
Calcic oxide (CaO) . . .	0·69
Magnesic oxide (MgO) . Potassic (OK_2) and sodic oxides (ONa_2)	traces.
	99·90

Clays containing calcic carbonate, indicated by their effervescing on addition of hydrochloric acid, are called *marls*. These silicates are used in the manufacture of earthenware and porcelain, of glass, of alum, &c.

Tests for Aluminium Compounds.—Potassic or sodic hydrate, added drop by drop, causes a white gelatinous precipitate of aluminic hydrate (Al_2Ho_6), which is soluble in excess of the reagent. Ammonic hydrate causes the same precipitate, but does not redissolve it when in excess. Ammonic sulphide also precipitates the hydrate with simultaneous evolution of sulphuretted hydrogen (SH_2). Heated on charcoal with cobaltous nitrate, aluminic compounds give a blue mass.

Platinum—Pt.

Atomic weight, 197·12. Atomicity, IV. and II. Sp. gr., 21·5.

Special Notice.—This metal was discovered in 1741, by Wood, an assay-master in Jamaica. Its name signifies *little silver.*

Occurrence in Nature.—Platinum occurs in nature in admixture, as alloys, with the metals rhodium, palladium, iridium, ruthenium, osmium, gold, silver, copper, iron, and lead. The localities where it occurs are the Ural Mountains in Russia, Mexico, Brazil, California, and Australia.

Preparation.—Wollaston has the honour of having invented the process by which platinum was first extracted. He treated the ore with *aqua regia*, which removes the most easily oxidizable metals first, and on continuing the digestion with application of heat the platinum dissolves:—

$$2Pt + \overbrace{\underset{\text{Hydrochloric acid.}}{8HCl} + \underset{\text{Nitric acid.}}{4NO_2Ho}}^{\text{Aqua regia.}} = \underset{\text{Platinic chloride.}}{2PtCl_4} + \underset{\text{Nitric oxide.}}{\left\{\begin{matrix}NO\\NO\end{matrix}\right.} + \underset{\text{Nitric peroxide.}}{N_2O_4} + \underset{\text{Water.}}{6OH_2}.$$

When complete solution has occurred, or when nothing soluble remains, the liquid is decanted and mixed with a solution of ammonic chloride; this causes the separation of a double chloride of ammonium and platinum ($PtCl_4,2AmCl$) in yellow crystals. The crystals are dried and ignited, the ignition decomposes them, ammonic chloride and chlorine being evolved, and platinum in a state known as *spongy platinum* remains :—

$PtCl_4,2AmCl$	=	$2AmCl$	+	$2Cl_2$	+	Pt
Double chloride of Pt and NH_4.		Ammonic chloride.		Chlorine.		Platinum.

The platinum is washed, put into a brass mould, and pressed ; the mass becomes compact, and assumes a metallic appearance. It is then heated in a wind furnace, and hammered on its ends into an ingot.

MM. Deville and Debray fuse the platinum in a crucible made of calcic oxide, by means of the oxyhydrogen flame.

Properties.—Platinum is a white, lustrous metal, about as hard as copper. It is equal to iron in ductility, and very tenacious. It is infusible at all known forge-heats, but can be converted into vapour by the oxyhydrogen blowpipe flame or the heat of the electric arc. Its sp. gr. is the same as those of osmium and iridium, these three metals being the heaviest substances known. It is the least expansive of the metals under heat, its co-efficient of expansion being nearly the same as that of glass, and hence it can be fused into that substance, a property which often admits of application in chemical apparatus. It is not oxidized at any temperature in air ; no single acid has any effect on it, but it is dissolved, though only slowly, by *aqua regia:* heated to redness with nitrates or hydrates, it is corroded. Compounds of easily reducible metals ought never to be heated in platinum crucibles, as they unite with the platinum, forming a readily fusible alloy, and so destroy the crucible. Spongy platinum becomes red-hot in a jet of hydrogen gas.

Uses.—Platinum is used for the manufacture of

Tests for Platinum Compounds.—Platinic compounds give a yellow precipitate of double chloride ($PtCl_4,2KCl$ or $PtCl_4,2AmCl$) with potassic or ammonic chloride. (See Platinic Chloride.) Sulphuretted hydrogen precipitates the black sulphide (PtS_2), insoluble in nitric acid. Ammonic sulphide gives the same precipitate, and redissolves it when present in large excess. The double chloride with ammonium is decomposed by heat, yielding metallic platinum.

Lead—Pb.

Atomic weight, 207. Atomicity, IV. and II., also occasionally a pseudo-triad. Sp. gr., 11·445.

Occurrence in Nature.—Lead occurs in nature as sulphide (PbS), known as *galena* and *lead glance*. It is also found occasionally as carbonate ($COPbo''$), known as *lead spar* and *white lead ore*. England and Spain yield it in largest quantities.

Preparation.—The sulphide is separated mechanically from the other substances with which it is mixed, and is then roasted in a reverberatory furnace, through which air passes; part of the sulphur burns off, part of the lead is converted into oxide, and part into sulphate :—

$$\text{(1.)}\quad \underset{\substack{\text{Plumbic}\\ \text{sulphide.}}}{2PbS} + \underset{\text{Oxygen.}}{3O_2} = \underset{\text{Plumbic oxide.}}{2PbO} + \underset{\substack{\text{Sulphurous}\\ \text{anhydride.}}}{2SO_2}.$$

$$\text{(2.)}\quad PbS + 2O_2 = \underset{\text{Plumbic sulphate.}}{SO_2Pbo''}.$$

When this oxidation has proceeded far enough, the air current is cut off, the furnace doors are closed, and the temperature raised, the sulphate and oxide react upon the sulphide, sulphurous anhydride is evolved, and metallic lead remains :—

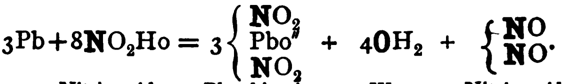
$$(1.)\ \mathbf{Pb}S + 2\mathbf{Pb}O = 3Pb + \mathbf{S}O_2.$$

Plumbic sulphide. Plumbic oxide. Lead. Sulphurous anhydride.

$$(2.)\ \mathbf{Pb}S + \mathbf{S}O_2Pbo'' = 2Pb + 2\mathbf{S}O_2.$$

Plumbic sulphate.

Properties.—Lead is of a bluish white colour, and so soft as to leave marks when rubbed on paper. It is neither very malleable nor very ductile. It melts at 334° C., and contracts on cooling; it may be crystallized in cubic or octahedral crystals. It speedily tarnishes in air, and also in *pure* water if exposed to the air at same time; a white film makes its appearance, covering the whole of the lead. This deposit is chiefly due to the action of the carbonic anhydride in the water, and consists of a combination of plumbic carbonate and hydrate, $\mathbf{C}OPbo'',\mathbf{Pb}Ho_2$. When heated in air it oxidizes rapidly, giving off white fumes of oxide; and if in a finely divided state, as obtained by igniting its tartrate in a narrow tube to which air has no access, it takes fire in air spontaneously. Sulphuric and hydrochloric acids have very little action on it, even when boiling; but nitric acid dissolves it readily, evolving nitric oxide:—



$$3Pb+8\mathbf{N}O_2Ho = 3\left\{\begin{matrix}\mathbf{N}O_2\\ Pbo''\\ \mathbf{N}O_2\end{matrix}\right. + 4OH_2 + \left\{\begin{matrix}\mathbf{N}O\\ \mathbf{N}O\end{matrix}\right..$$

Nitric acid. Plumbic nitrate. Water. Nitric oxide.

Lead is nearly pure as met with in commerce.

Uses.—Lead is extensively used in the form of sheets in roofing, in making sulphuric acid chambers, and in forming water-pipes, cisterns, &c. Alloyed with arsenic sufficient to facilitate its assuming the globular form, it is used in making shot; with antimony and tin it forms *type-metal*, an alloy which expands on cooling. *Fusible metal*, *pewter*, *Britannia metal*, *solder*, &c., are other alloys containing lead.

Principal Compounds.

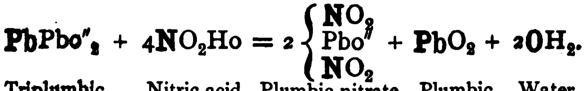
Plumbic Oxide (protoxide of lead, litharge, massicot, lead monoxide), PbO. This oxide is obtained by heating lead in a current of air :—

$$2Pb + O_2 = 2PbO.$$

It is of a yellow colour. It is soluble in sodic hydrate, and the solution, on exposure to air, absorbs carbonic anhydride and deposits the oxide in dodecahedral crystals. It fuses at a red heat, and crystallizes in scales on cooling. It gradually absorbs carbonic anhydride from the air. It is used as a flux in assaying, in making optical glass, and is a constituent of the glaze of earthenware. It dissolves in acids, forming plumbic salts.

A plumbous oxide ($'Pb'_2O$) of a black colour is also known.

Plumbic Peroxide (peroxide, binoxide, puce-coloured or brown oxide of lead, lead dioxide), PbO_2.—This oxide is obtained by treating *red-lead* (see below) with nitric acid; the plumbic oxide contained in it is dissolved, and plumbic peroxide remains, and is separated by filtering and washing :—



$$PbPbo''_2 + 4NO_2Ho = 2\left\{\begin{matrix} NO_2 \\ Pbo'' \\ NO_2 \end{matrix}\right. + PbO_2 + 2OH_2.$$

Triplumbic tetroxide (red-lead).	Nitric acid.	Plumbic nitrate.	Plumbic peroxide.	Water.

Plumbic peroxide is a brown powder. At a red heat it is converted into plumbic oxide and oxygen:— $2\mathbf{Pb}O_2 = 2\mathbf{Pb}O + O_2$. Hydrochloric acid converts it into plumbic chloride, with evolution of chlorine:— $\mathbf{Pb}O_2 + 4HCl = \mathbf{Pb}Cl_2 + 2\mathbf{O}H_2 + Cl_2$. When heated with sulphuric acid it forms plumbic sulphate, oxygen being liberated:—$2\mathbf{Pb}O_2 + 2\mathbf{S}O_2Ho_2 = 2\mathbf{S}O_2Pbo'' + 2\mathbf{O}H_2 + O_2$. Plumbic peroxide is used to separate sulphurous anhydride from gaseous mixtures:—

$$\mathbf{Pb}O_2 + \mathbf{S}O_2 = \mathbf{S}O_2Pbo''.$$

It acts as a feeble anhydride, forming salts called *plumbates.*

Red-lead.—Three compounds are known in commerce as *red-lead*, *minium*, or *red oxide of lead;* these are diplumbic trioxide ($\mathbf{Pb}OPbo''$), triplumbic tetroxide ($\mathbf{Pb}Pbo''_2$), and tetraplumbic pentoxide ($\mathbf{Pb}Pbo''[Pb_2O_2]''$).

They are prepared by exposing for a long time to a dull red heat the plumbic oxide, which has not been fused. The result is a brilliant red powder. By the action of acids they are converted into solutions of plumbic salts and plumbic peroxide. They are used in painting, and sold as a substitute for *vermilion;* but their chief use is in glass-making, both *flint* and *optical.* At a high temperature they are decomposed, oxygen being given off.

Plumbic Sulphide (galena, lead glance), $\mathbf{Pb}S$.—This is the most abundant ore of lead; it is found crystallized in cubes, very metallic-looking, and of a deep lead colour. The hydrated sulphide is obtained by passing sulphuretted hydrogen through a solution of a plumbic salt. It is decomposed by nitric acid, *aqua regia*, and boiling hydrochloric acid.

Plumbic Chloride (chloride of lead), $\mathbf{Pb}Cl_2$.—

If hydrochloric acid or sodic chloride be added to a solution of plumbic nitrate, a precipitate of plumbic chloride is obtained as a heavy white powder:—

$$\left\{\begin{matrix}NO_2\\ Pbo''\\ NO_2\end{matrix}\right. + 2HCl = PbCl_2 + 2NO_2Ho.$$

Plumbic nitrate. Hydrochloric acid. Plumbic chloride. Nitric acid.

It dissolves in 33 times its weight of boiling water, and separates in six-sided needles on cooling. It is easily fusible, and is volatile at a high temperature.

Plumbic Sulphate (sulphate of lead, lead vitriol), SO_2Pbo''.—This salt occurs in nature, and is obtained by adding sulphuric acid to a solution of a plumbic salt. It forms a white powder, slightly soluble in nitric acid, but insoluble in pure water. It is formed in the extraction of lead and in the preparation of aluminic acetate. It is reduced to metallic lead by heating in the reducing blowpipe flame.

Plumbic Nitrate (nitrate of lead, plumbic dinitrate), $\left\{\begin{matrix}NO_2\\ Pbo''\\ NO_2\end{matrix}\right.$.—This salt is prepared by dissolving plumbic oxide or lead in nitric acid:—

$$(1)\quad PbO + 2NO_2Ho = \left\{\begin{matrix}NO_2\\ Pbo''\\ NO_2\end{matrix}\right. + OH_2.$$

Plumbic oxide. Nitric acid. Plumbic nitrate. Water.

(2) See page 264.

It is soluble in 8 times its weight of water, and crystallizes in regular octahedra; heated to redness it is decomposed as follows:—

$$2\left\{\begin{matrix}NO_2\\ Pbo''\\ NO_2\end{matrix}\right. = \underset{\text{Plumbic oxide.}}{2PbO} + \underset{\text{Nitric peroxide.}}{2N_2O_4} + \underset{\text{Oxygen.}}{O_2.}$$

It is used in preparing plumbic chromate (chrome yellow).

Plumbic Carbonate (carbonate of lead, white-lead, lead spar, white-lead ore), $COPbo''$.—This salt is sometimes found crystallized in long white needles with other lead ores. It may be prepared by adding in the cold a solution of sodic carbonate to a solution of plumbic acetate. It is, however, largely manufactured for painting by the following process, known as the *Dutch method*:—Into a great number of earthen jars, arranged side by side, a little vinegar (weak acetic acid) is poured. These jars stand on manure or tan; to each jar a roll of sheet lead is adapted, or bars cast into the shape of a portcullis are placed, so as not to touch the vinegar or project above the jar. The jars are loosely covered with boards, and a layer of tan or manure is placed upon the top. Another row of pots is then placed upon the top of these, covered in like manner. By this arrangement a pile of many rows is constructed. After the lapse of some days the sheets or bars of lead are taken out of the pots, when it is found that they are in great part converted into the carbonate, which is made ready for use by washing and grinding. The nature of the change is explained as follows:—the heat of the decomposing manure, or tan, causes the acetic acid to volatilize and attack the lead, forming plumbic acetate, which is then decomposed by the carbonic anhydride of the decomposing manure:—

$$\underset{\substack{\text{Plumbic}\\\text{acetate.}}}{\left\{\begin{array}{l}\mathrm{C(Me)O}\\\mathrm{Pbo''}\\\mathrm{C(Me)O}\end{array}\right.} + \underset{\substack{\text{Carbonic}\\\text{anhydride.}}}{\mathrm{CO}} + \underset{\text{Water.}}{\mathrm{OH_2}} = \underset{\substack{\text{Plumbic}\\\text{carbonate.}}}{\mathrm{COPbo''}} + 2\underset{\substack{\text{Acetic}\\\text{acid.}}}{\left\{\begin{array}{l}\mathrm{CH_3}\\\mathrm{COHo}\end{array}\right.}.$$

The acetic acid thus set free acts on another portion of lead, and so the process is continuous. The carbonate, although very insoluble in pure water, is slightly soluble in water containing carbonic anhydride, which is nearly always the case. If the water were quite pure it would be contaminated with lead, and in course of time the whole of the lead would be dissolved. As all lead compounds are very poisonous, pure water, or water containing much free carbonic anhydride, cannot therefore be kept in leaden cisterns, or passed through leaden pipes. In common river and spring water, however, a little calcic sulphate is almost always present; this forms a thin film of the insoluble plumbic sulphate, which prevents further action. It is for this reason that leaden cisterns may generally be used without danger.

Plumbic Chromate (chromate of lead, chrome yellow, lead chromate), $\mathrm{CrO_2Pbo''}$.—This salt is obtained as a yellow precipitate by adding a solution of potassic chromate, or bichromate, to a solution of plumbic acetate or nitrate:—

$$\underset{\substack{\text{Plumbic}\\\text{nitrate.}}}{\left\{\begin{array}{l}\mathrm{NO_2}\\\mathrm{Pbo''}\\\mathrm{NO_2}\end{array}\right.} + \underset{\substack{\text{Potassic}\\\text{chromate.}}}{\mathrm{CrO_2Ko_2}} = \underset{\text{Plumbic chromate.}}{\mathrm{CrO_2Pbo''}} + \underset{\text{Potassic nitrate.}}{2\mathrm{NO_2Ko}}.$$

When boiled with calcic hydrate, plumbic chromate becomes of a scarlet colour, owing to its being converted into diplumbic chromate:—

$$2CrO_2Pbo'' + CaHo_2 = CrOPbo'_2 + CrO_2Cao'' + OH_2$$

Calcic hydrate. Diplumbic chromate. Calcic chromate. Water.

It is used in dyeing yellows and oranges, and for supplying oxygen in certain cases of combustion analysis.

Tests for Lead Compounds.—The soluble salts redden litmus. Sulphuric acid and solutions of sulphates precipitate the white sulphate (SO_2Pbo''), almost insoluble in water and acids. Hydrochloric acid and chlorides precipitate from concentrated and moderately dilute solutions the white crystalline chloride ($PbCl_2$), which redissolves on warming. Sulphuretted hydrogen converts all compounds of lead, whether soluble or insoluble, into plumbic sulphide (PbS), insoluble in potassic and sodic hydrates. Potassic chromate precipitates the yellow chromate (CrO_2Pbo''); potassic iodide, the yellow iodide PbI_2). Heated in the reducing blowpipe flame on charcoal with sodic carbonate, they yield metallic globules, which make a black mark when rubbed on paper.

Tin—Sn.

Atomic weight, 118. Atomicity, IV. and II. Sp. gr., 7·28.

Special Notice.—Tin was known to the ancients. The Phœnicians are said to have traded with the ancient Britons for tin.

Occurrence in Nature.—Tin occurs in nature principally as stannic oxide or anhydride (SnO_2), commonly called *tinstone* or *cassiterite* (from *Cassiterides,* the tin islands,—now the Scilly Isles, off the coast of Cornwall). The chief sources of the ore of this metal are the mines of Cornwall, Saxony, Bohemia, and Malacca. It is sometimes found as an alluvial deposit in the beds of many small rivers, and is then known as *stream tin;* this variety, being very pure, is the most valuable form in which the ore occurs.

Preparation.—The ore is first subjected to the mechanical operations of picking, stamping, and washing; it is then roasted on the bed of a reverberatory furnace in order to expel volatile impurities, such as arsenic and sulphur, which it usually contains. The resulting slightly purified mass is then treated with water, which dissolves out any cupric sulphate formed by the oxidation of the variable amount of *copper pyrites* usually mixed with the ore. The ore is then ready for reduction, and this is accomplished by heating it along with about one-eighth its weight of anthracite coal, or charcoal and a little lime, in a low-crowned reverberatory furnace. The lime forms a fusible slag by union with the siliceous matter yet remaining in the ore, and the carbon of the coal reduces the stannic oxide to metallic tin, with formation of carbonic oxide:—

$$\underset{\text{Stannic oxide.}}{SnO_2} + \underset{\text{Carbon.}}{2C} = \underset{\text{Carbonic oxide.}}{2CO} + \underset{\text{Tin.}}{Sn.}$$

When the reduction is complete the temperature of the furnace is raised very high, so as to melt the slag completely; the melted metal sinks beneath it, and is drawn off into iron pans, and then cast into ingots. The metal so obtained is not pure; it is therefore remelted; the purer tin, owing to its lower fusibility, melts first and runs off, and is further purified by thrusting pieces of wood soaked in water into it while melted.

Properties.—Tin is a brilliant white metal. It is not very hard nor tenacious, but is very malleable. If melted and allowed to cool slowly it may be obtained in crystals. At 100° it becomes moderately ductile. It is not much affected at ordinary temperatures by exposure to the action of water or air; when raised to

Compounds of Tin.

Stannic Compounds (Compounds of Tetrad Tin).

Stannic Oxide (binoxide of tin), SnO_2.—This oxide is, as has already been stated, the principal natural compound of tin. It sometimes occurs in square prisms, having a brown tinge due to ferric or manganic oxide; the crystals are hard enough to scratch glass. Acids have no action on it, but if fused with potassic or sodic hydrate, a compound is obtained which is soluble in water; it is, in fact, the potassic or sodic salt of stannic acid. The oxide produced by drying metastannic acid at 100° is called *putty powder*, and is used for polishing plate and preparing white opaque enamels.

Stannic Acid (hydrated binoxide of tin), $SnOHo_2$.—This acid is prepared by adding calcic carbonate to an excess of a solution of stannic chloride:—

$$\underset{\text{Stannic chloride.}}{SnCl_4} + \underset{\text{Calcic carbonate.}}{2COCao''} + \underset{\text{Water.}}{OH_2} = \underset{\text{Stannic acid.}}{SnOHo_2} + \underset{\text{Calcic chloride.}}{2CaCl_2} + \underset{\text{Carbonic anhydride.}}{2CO_2}$$

The acid falls as a gelatinous precipitate, and can be readily purified by washing with water.

Sodic Stannate (stannate of soda), $SnONao_2, 4OH_2$.—This is the most important salt of stannic acid, and is prepared by fusing stannic oxide with excess of sodic hydrate dissolving the fused mass in water, and evaporating the solution; the salt then crystallizes in six-sided tables. Like some of the salts of calcium, it is more soluble in cold water than in hot. It is now prepared on the large scale for use as a mordant by the dyer and calico-printer.

Metastannic Acid, $Sn_5O_5Ho_{10}$.—To prepare this acid it is only necessary to act upon metallic tin with nitric acid; the nitric acid gives up oxygen to the tin, large quantities of nitric peroxide are evolved, and the metastannic acid is produced as a white crystalline substance insoluble in water. It reddens litmus, and becomes converted into stannic oxide by drying at 100°. Metastannic acid is insoluble in nitric and hydrochloric acids; it combines with the latter, however, and the compound, though insoluble in the excess of acid, is soluble in water. The only soluble salts of this acid are those of potassium and sodium.

Stannic Chloride (tetrachloride, bichloride, or nitromuriate of tin), $\mathrm{SnCl_4}$.—This important salt is prepared (1) by heating tin in a tube through which chlorine passes:—$\mathrm{Sn} + 2\mathrm{Cl_2} = \mathrm{SnCl_4}$; (2) by distilling an intimate mixture of mercuric chloride (4 parts) and metallic tin 1 part:—

$$2\mathrm{HgCl_2} + \mathrm{Sn} = \mathrm{SnCl_4} + 2\mathrm{Hg};$$

(3) by dissolving tin in a mixture of nitric and hydrochloric acids (commonly called *nitromuriatic acid* or *aqua regia*):—

$$\mathrm{Sn} + 4\mathrm{HCl} + 2\mathrm{NO_2Ho} = \mathrm{SnCl_4} + 3\mathrm{OH_2} + \left\{\begin{matrix}\mathrm{NO}\\ \mathrm{O}\end{matrix}\right. .$$

having the power of uniting with colouring matter, fixes the latter on the cloth when worked in the dye-bath.

Stannic Sulphide (bisulphide of tin), $\mathbf{Sn}S_2$.—This compound is obtained in yellow flakes known as *mosaic gold* by heating to low redness in a glass flask a mixture of 12 parts tin, 6 parts mercury (these two amalgamated together before mixing with the other ingredients), 6 parts of ammonic chloride, and 7 parts of flowers of sulphur; mercurous chloride and mercuric sulphide sublime together with ammonic chloride, and the stannic sulphide remains. It is used in the arts in making imitation bronzes. The hydrated sulphide is precipitated as a dingy yellow powder on passing sulphuretted hydrogen through a solution of stannic salt.

Stannous Compounds (Compounds of Diad Tin).

Stannous Oxide (protoxide of tin), $\mathbf{Sn}O$.—By ignition of stannous hydrate in an atmosphere of hydrogen or carbonic anhydride this compound is obtained as a black powder. If heated with acids it dissolves, though slowly; in the cold very little action takes place: the solutions are stannous salts. It may be obtained in black crystalline needles by boiling stannous hydrate with a dilute solution of potassic hydrate in quantity insufficient to dissolve it.

Stannous Hydrate (hydrate of protoxide of tin), $\mathbf{Sn}Ho_2$.—This substance is obtained as a white precipitate by adding stannous chloride solution to a solution of sodic carbonate:—

$$\underset{\substack{\text{Sodic}\\\text{carbonate.}}}{\mathbf{C}ONao_2} + \underset{\substack{\text{Stannous}\\\text{chloride.}}}{\mathbf{Sn}Cl_2} + \underset{\text{Water.}}{\mathbf{O}H_2} = \underset{\substack{\text{Stannous}\\\text{hydrate.}}}{\mathbf{Sn}Ho_2} + \underset{\substack{\text{Sodic}\\\text{chloride.}}}{2NaCl} + \underset{\substack{\text{Carbonic}\\\text{anhydride.}}}{\mathbf{C}O_2}.$$

Stannous hydrate is soluble in potassic hydrate and in acids.

Stannous Chloride (protochloride of tin), $SnCl_2$.—The ordinary method of preparing this salt is by dissolving tin in hydrochloric acid :—

$$Sn + 2HCl = SnCl_2 + H_2.$$

As thus prepared it crystallizes in prismatic needles whose composition is $SnCl_2,2OH_2$. The anhydrous chloride is obtained when a mixture of equal weights of tin and mercuric chloride are distilled; mercury passes off in vapour, and the chloride is left behind as a vitreous mass of a brilliant grey colour. It distils when heated to redness :—

$$HgCl_2 + Sn = SnCl_2 + Hg.$$

Butter of tin is an old name for the anhydrous chloride. Aqueous solutions of this salt, under the names of *tin solutions* and *tin spirits*, are extensively used by the dyer and calico-printer as mordants. In some cases a small quantity of oxalic acid is added to the tin and hydrochloric acid, in preparing these solutions; the resulting liquors are then called *ox tin* or *oxalic of tin*. The solutions do not, however, contain any oxalic acid, but another organic acid called glycolic acid, which is formed by the action of the nascent hydrogen upon the oxalic acid. As already stated under stannic chloride, the efficacy of a mordant depends upon its power of fixing the colouring matter upon the fibre of the cloth. An experiment illustrative of this, as well as of the use of stannous chloride in dyeing scarlets, may be made as follows :—Make an infusion of cochineal by boiling up a small quantity of the powdered dry insects with water, and boil in this solution of colouring matter two pieces of clean white flannel or Berlin wool of the same size; after boiling *both* a few minutes, remove one of the pieces, add a few drops of stannous chloride to the infusion, and boil again; then remove the other piece. On now comparing the pieces, it will be found that the piece dyed without any mordant is of a purplish tint, while the one with the mordant is of a bright scarlet; and further, if the two be washed with water repeatedly, it will be found that the former will lose almost all its colour, while the latter will only suffer a slight diminution in shade. The addition of a little cream of tartar and a chip of young fustic will improve the colour.

Owing to its tendency to unite with chlorine and oxygen, stannous chloride is a powerful reducing agent; examples of its reducing action are the production of gold, silver, and mercury in the metallic state, when a solution of it is added to solutions of salts of these metals; the reaction with gold salts may be taken as typical:—

$$\underset{\text{Auric chloride.}}{2AuCl_3} + \underset{\text{Stannous chloride.}}{3SnCl_2} = \underset{\text{Gold.}}{Au_2} + \underset{\text{Stannic chloride.}}{3SnCl_4}.$$

Stannous Sulphide (protosulphide of tin), **SnS**.—This substance is produced by mixing tin and sulphur in the proportions of their atomic weights, and heating the mixture in a crucible; as thus prepared it is of a bluish grey colour. The hydrated form of it may be obtained by passing sulphuretted hydrogen through a solution of stannous salt; it separates as a chocolate-brown precipitate. It is soluble in strong hydrochloric acid, with evolution of sulphuretted hydrogen.

Tests for Compounds of Tin.—Stannic Salts.—Potassic and sodic hydrates precipitate stannic acid ($SnOHo_2$) as a white powder, which an excess of the reagent re-dissolves. Ammonic hydrate, and potassic, sodic, and ammonic carbonates produce the same precipitate (the three latter with escape of carbonic anhydride), but do not re-dissolve it when added in excess. Sulphuretted hydrogen and ammonic sulphide precipitate the dingy yellow hydrated sulphide (SnS_2), which is soluble in potassic, sodic, and ammonic hydrates and sulphides.

Stannous Salts.—Potassic and sodic hydrates cause a white precipitate of stannous hydrate, which redissolves in excess of the reagent, and on boiling the solution stannous oxide is deposited in black crystals. Ammonic hydrate, and potassic, sodic, and ammonic carbonates produce the some precipitate (carbonic anhydride escaping at the same time in the latter three cases), but do not re-dissolve it when added in excess. Sulphuretted hydrogen and ammonic sulphide precipitate the characteristic chocolate-brown hydrated sulphide (SnS), which is soluble in excess of ammonic sulphide, and in potassic and sodic sulphides. Auric chloride in dilute solution and used in excess gives a brown precipitate of gold, owing to the reducing action of stannous salts; but if the tin salt be in excess, a beautiful purple precipitate, known as *purple of Cassius*, results.

Both stannic and stannous salts, when heated with sodic carbonate on charcoal before the blowpipe, yield globules of metallic tin.

HEXAD METALS.

Chromium—Cr.

Atomic weight, 52·5. Molecular weight, unknown. Molecular volume, unknown. Atomicity, VI., IV., and II.; also a pseudo-triad, and occasionally a pseudo-octad. Sp. gr., 6.

Special Notice.—Chromium was discovered in 1797 by Vauquelin, in plumbic chromate (CrO_2Pbo''), known as *crocoisite* or *red-lead ore.*

Occurrence in Nature.—Chromium compounds are not very abundant in nature. It occurs principally as *chrome-iron ore* or *chrome ironstone*, which is dichromic-ferrous tetroxide$\left(\left\{\begin{matrix}CrO\\CrO\end{matrix}\right.Feo''\right)$; this occurs in the Shetland Isles, Sweden, America, Australia, and New Zealand.

Preparation.—The metal is procured by mixing chromic oxide with one-fifth of its weight of charcoal powder, and placing the mixture in a crucible lined with charcoal. The crucible is then placed in a furnace, and raised to the very highest heat. The chromic oxide is reduced by the charcoal, and the metal is obtained as a greyish white brittle substance, and very difficult of fusion:—

$$\underset{\text{Chromic oxide.}}{\left\{\begin{matrix}CrO\\CrO\end{matrix}\right.O} \quad + \quad \underset{\text{Carbon.}}{3C} \quad = \quad \underset{\text{Chromium.}}{2Cr} \quad + \quad \underset{\text{Carbonic oxide.}}{3CO.}$$

Properties.—The metal has been seen by very few persons. It crystallizes in a form belonging to the regular system. It is oxidized by nitric acid with such violence when in a pulverulent form as to become red-hot occasionally. It burns brilliantly when heated in air. Hydrochloric and sulphuric acids dissolve it with evolution of hydrogen.

Uses.—No use has as yet been found for metallic chromium. It is probable, however, than when its properties have been thoroughly investigated it will be applied in the useful arts.

Principle Compounds.—Owing to its varying active atomicity chromium forms numerous compounds.

Chromous Compounds.

Chromous Oxide (protoxide of chromium), **CrO**.—This oxide has never yet been obtained in the anhydrous form; chromous hydrate (**$CrHo_2$**) is, however, obtainable by adding potassic hydrate to a solution of chromous chloride:—

$$\underset{\text{Chromous chloride.}}{CrCl_2} + \underset{\text{Potassic hydrate.}}{2KHo} = \underset{\text{Chromous hydrate.}}{CrHo_2} + \underset{\text{Potassic chloride.}}{2KCl.}$$

It forms a dark brown powder.

Chromous Chloride (protochloride of chromium), **$CrCl_2$**, is obtained by heating to redness chromic chloride $\left(\left\{\begin{matrix}CrCl_3\\CrCl_3\end{matrix}\right.\right)$ in a porcelain tube, through which a current of pure dry hydrogen is passing; hydrochloric acid is produced, and a white mass of chromous chloride is obtained, which dissolves in water with a great increase of temperature, yielding a blue solution, which absorbs oxygen by exposure to the air, and passes into a deep green colour.

$$\underset{\text{Chromic chloride.}}{\left\{\begin{matrix}CrCl_3\\CrCl_3\end{matrix}\right.} + \underset{\text{Hydrogen.}}{H_2} = \underset{\text{Chromous chloride.}}{2CrCl_2} + \underset{\text{Hydrochloric acid.}}{2HCl}$$

Chromous chloride is one of the most powerful reducing agents known.

Chromic Compounds.

Chromic Oxide (sesquioxide of chromium), $\left\{\begin{matrix}CrO\\CrO\end{matrix}\right.O$.—If chromic hydrate be heated to redness it is decomposed, water (as steam) is evolved, and chromic oxide remains:—

$$\underset{\text{Chromic hydrate.}}{\left\{\begin{matrix}CrHo_3\\CrHo_3\end{matrix}\right.} = \underset{\text{Chromic oxide.}}{\left\{\begin{matrix}CrO\\CrO\end{matrix}\right.O} + \underset{\text{Water.}}{3OH_2.}$$

It is a green powder when so obtained, and is insoluble in acids. It is used as a green pigment in the formation of a pink enamel colour, and in calico-printing. United with ferrous oxide it forms *chrome-iron ore*, the principal ore of the metal, which is often found crystallized in octahedra, like *magnetic iron-ore.* It is infusible in the furnace, and when hot absorbs oxygen from the air, forming chromic anhydride; and if some oxide, hydrate, or carbonate be present, a chromate is formed.

Chromic Hydrate (hydrated sesquioxide of chromium), $\left\{\begin{matrix}\mathbf{Cr}\mathrm{Ho_3}\\ \mathbf{Cr}\mathrm{Ho_3}\end{matrix}\right.$.—This substance is produced by adding ammonic hydrate to a solution of chromic chloride:—

$$\underset{\text{Chromic chloride.}}{\left\{\begin{matrix}\mathbf{Cr}\mathrm{Cl_3}\\ \mathbf{Cr}\mathrm{Cl_3}\end{matrix}\right.} + \underset{\text{Ammonic hydrate.}}{6\mathrm{AmHo}} = \underset{\text{Chromic hydrate.}}{\left\{\begin{matrix}\mathbf{Cr}\mathrm{Ho_3}\\ \mathbf{Cr}\mathrm{Ho_3}\end{matrix}\right.} + \underset{\text{Ammonic chloride.}}{6\mathrm{AmCl}}.$$

It forms a dark green powder, very bulky and gelatinous. It dissolves in acids, forming chromic salts.

Chromic Anhydride (chromic acid), $\mathbf{Cr}^{\mathrm{vi}}\mathrm{O_3}$.—Chromic anhydride is prepared by adding 5 measures of sulphuric acid to 4 measures of a cold saturated solution of dipotassic dichromate, and drying the beautiful crimson crystals so obtained at a gentle heat on a brick or a porous tile:—

$$\underset{\text{Dipotassic dichromate.}}{\left\{\begin{matrix}\mathbf{Cr}\mathrm{O_2Ko}\\ \mathrm{O}\\ \mathbf{Cr}\mathrm{O_2Ko}\end{matrix}\right.} + \underset{\text{Sulphuric acid.}}{\mathbf{S}\mathrm{O_2Ho_2}} = \underset{\text{Potassic sulphate.}}{\mathbf{S}\mathrm{O_2Ko_2}} + \underset{\text{Water.}}{\mathbf{O}\mathrm{H_2}} + \underset{\text{Chromic anhydride.}}{2\mathbf{Cr}\mathrm{O_3}}.$$

When heated to 204° it melts to a black liquid, and at a higher temperature it is converted into chromic oxide and oxygen with incandescence:—

$$\underset{\text{Chromic anhydride.}}{4\mathbf{Cr}\mathrm{O_3}} = \underset{\text{Chromic oxide.}}{2\mathbf{Cr}_2\mathrm{O_3}} + \underset{\text{Oxygen.}}{3\mathrm{O_2}}.$$

It is deliquescent, absorbing moisture from the air, with which it combines, and forms chromic acid (CrO_2Ho_2). Heated with hydrochloric acid, chromic chloride, water, and chlorine are produced :—

$$\underset{\text{Chromic anhydride.}}{2CrO_3} + \underset{\text{Hydrochloric acid.}}{12HCl} = \underset{\text{Chromic chloride.}}{\left\{\begin{matrix}CrCl_3\\ CrCl_3\end{matrix}\right.} + \underset{\text{Water.}}{6OH_2} + \underset{\text{Chlorine.}}{3Cl_2}.$$

Chromic Acid, CrO_2Ho_2.—If chromic anhydride be added to water, it combines with it, forming chromic acid :—

$$\underset{\text{Chromic anhydride.}}{CrO_3} + \underset{\text{Water.}}{OH_2} = \underset{\text{Chromic acid.}}{CrO_2HO_2}.$$

It is a powerful oxidizing agent; for instance, if thrown into alcohol it sets it on fire, so powerful is its oxidizing action. The most important of its salts are the potassium ones.

Dipotassic Chromate (chromate of potash, neutral or yellow chromate of potash), Cr_2Ko_2, is the source from which all the chromic preparations are obtained. It is made from chromic iron ore. The ore is reduced to powder, calcined and heated for a long time in a furnace with crude potassic carbonate and a little potassic nitrate. The product is treated with water, and yields a yellow solution, which deposits, on evaporation, crystals of the same colour :—

$$\underset{\text{Chrome-iron ore.}}{2\left\{\begin{matrix}CrO\\ CrO\end{matrix}\right.Feo''} + \underset{\text{Potassic carbonate.}}{4COKo_2} + \underset{\text{Oxygen (from } \textit{nitre} \text{ and air).}}{5O_2} = \underset{\text{Potassic chromate.}}{4CrO_2Ko_2} + \underset{\text{Ferric oxide.}}{2Fe_2O_3}$$

$$+ \underset{\text{Carbonic anhydride.}}{4CO_2}.$$

Dipotassic Dichromate (bichromate or red chromate of potash), $\left\{\begin{matrix}\mathrm{CrO_2Ko}\\ \mathrm{O}\\ \mathrm{CrO_2Ko}\end{matrix}\right.$—This salt is prepared by adding sulphuric acid to a solution of the preceding salt; on evaporating the solution the salt is obtained in red four-sided tables of great beauty:—

$$\underset{\text{Potassic chromate.}}{2\mathrm{CrO_2Ko_2}} + \underset{\text{Sulphuric acid.}}{\mathrm{SO_2Ho_2}} = \underset{\text{Dipotassic dichromate.}}{\left\{\begin{matrix}\mathrm{CrO_2Ko}\\ \mathrm{O}\\ \mathrm{CrO_2Ko}\end{matrix}\right.} + \underset{\text{Potassic sulphate.}}{\mathrm{SO_2Ko_2}} + \underset{\text{Water.}}{\mathrm{OH_2}}.$$

It fuses below a red heat, and is decomposed at a higher temperature into chromic oxide, potassic chromate, and oxygen:—

$$\underset{\text{Dipotassic dichromate.}}{4\left\{\begin{matrix}\mathrm{CrO_2Ko}\\ \mathrm{O}\\ \mathrm{CrO_2Ko}\end{matrix}\right.} = \underset{\text{Dipotassic chromate.}}{4\mathrm{CrO_2Ko_2}} + \underset{\text{Chromic oxide.}}{2\left\{\begin{matrix}\mathrm{CrO}\\ \mathrm{CrO}\end{matrix}\right.\mathrm{O}} + \underset{\text{Oxygen.}}{3\mathrm{O_2}}.$$

It requires about 10 times its weight of water, at 16° C., to dissolve it. It is used as an oxidizing agent, and for preparing *chrome alum*. It is also much used in dyeing, and in the preparation of chrome yellow and orange; a saturated solution of it is used as the exciting fluid in the bottle battery, and it is in constant requisition in the laboratory.

Chromic Chloride (sesquichloride of chromium), $\left\{\begin{matrix}\mathrm{CrCl_3}\\ \mathrm{CrCl_3}\end{matrix}\right.$.—This compound may be obtained anhydrous by heating to redness in a porcelain tube,

It takes fire spontaneously in ammonia gas, and sets fire to sulphur and phosphorus when placed in contact with them.

Chrome Alum, Dipotassic-dichromic Tetrasulphate, $S_4O_8Ko_2(Cr_2O_6)^{vi}$, $24OH_2$.—If dipotassic dichromate be boiled with sulphuric acid, oxygen is given off, and this substance is formed; or if sulphurous anhydride be passed into it chrome alum is produced:—

(1) For equation see this mode of preparing oxygen.

$$(2)\ \left\{\begin{array}{l}CrO_2Ko\\ O\\ CrO_2Ko\end{array}\right. + 3SO_2 + SO_2Ho_2 = \underset{\text{Chrome alum.}}{S_4O_8Ko_2(Cr_2O_6)^{vi}} + OH_2.$$

It crystallizes in octahedra like common alum, but the crystals are almost black when seen by reflected light. In order to obtain fine crystals the solution should be set aside for a fortnight at least.

Tests for Chromium Compounds.—*Chromous salts* are never met with in ordinary analysis. They give a brown precipitate of chromous hydrate ($CrHo_2$) with potassic hydrate; ammonic sulphide gives a black sulphide (CrS). *Chromic salts* give, with potassic, sodic, and ammonic hydrates, a dirty green precipitate of chromic hydrate, $\left(\left\{\begin{array}{l}CrHO_3\\ CrHO_3\end{array}\right.\right)$; the two former redissolve the precipitate. Ammonic sulphide precipitates the hydrate as well, sulphuretted hydrogen being evolved at the same time. The chromic hydrate produced by these reagents, if fused with potassic nitrate and carbonate, is converted into potassic chromate; and if after the fusion it be dissolved in water, the solution gives a yellow precipitate of plumbic chromate (CrO_2Pbo'') with a soluble lead salt. Sodic chloro-hypochlorite (chloride of soda) also converts chromic salts into chromates. The *chromates* are distinguished by giving a yellow precipitate of plumbic chromate (CrO_2Pbo'') with a soluble lead salt, a pale yellow precipitate of baric chromate (CrO_2Bao'') with a soluble barium salt, and a purple-red precipitate of argentic chromate (CrO_2Ago_2) with argentic nitrate solution.

Manganese.—Mn.

Atomic weight, 55. Atomicity, VI., IV., and II.; also a pseudo-triad and a pseudo-octad. Sp. gr. 7 to 8.

Special Notice.—Manganese was discovered in 1774 by Gahn and Scheele.

Occurrence in Nature.—The compounds of manganese occur in nature pretty abundantly. Its most abundant and important ores are manganic oxide (MnO_2), known as *pyrolusite*; dimanganic trioxide $\left(\left\{\begin{matrix}MnO\\MnO\end{matrix}\right.O\right)$, known as *braunite;* dimanganic dioxydihydrate $\left(\left\{\begin{matrix}MnOHo\\MnOHo\end{matrix}\right.\right)$, and manganic oxide-hydrate (Mn_2,OH_2), known as *wad* or bog manganese. The localities for them are Devonshire, Cornwall, the Hartz Mountains, and Piedmont.

Preparation.—Manganese is best obtained by heating the carbonate ($COMno''$) made into a paste with oil and charcoal, in a crucible: the cover should be luted on and raised to the highest temperature of the forge for two hours.

$$\underset{\text{Manganous carbonate.}}{COMno''} + \underset{\text{Carbon.}}{2C} = \underset{\text{Manganese.}}{Mn} + \underset{\text{Carbonic oxide.}}{3CO.}$$

The metal is obtained as a button containing a little carbon.

Properties.—It is of a greyish white colour, resembling cast iron, hard and brittle and slightly magnetic. It oxidizes by exposure to the air, and decomposes water at the ordinary temperature:—

Mn	+	$2OH_2$	=	$MnHo_2$	+	H_2.
Manganese.		Water.		Manganous hydrate.		Hydrogen.

On account of these properties it has to be preserved either in hermetically sealed tubes or under napthha.

Uses.—Manganese is only known as a chemical curiosity; it is put to no use in the arts.

Principal Compounds.

Manganous Oxide (protoxide of manganese), MnO.—This oxide may be obtained by heating manganous carbonate in an atmosphere of hydrogen; carbonic oxide and water are given off, and manganous oxide remains:—

$COMno''$	+	H_2	=	MnO	+	CO	+	OH_2.
Manganous carbonate.		Hydrogen.		Manganous oxide.		Carbonic oxide.		Water,

It is an olive-green powder, which absorbs oxygen from the air. It dissolves in acids, forming manganous salts.

Manganous Hydrate (hydrated oxide of manganese), $MnHo_2$.—This substance is prepared by adding a solution of potassic hydrate to a solution of a manganous salt:—

$MnCl_2$	+	$2KHo$	=	$MnHo_2$	+	$2KCl$.
Manganous chloride.		Potassic hydrate.		Manganous hydrate.		Potassic chloride.

(1.)	MnO_2	+	$2C$	=	MnO	+	$2CO$.
	Manganic oxide.		Carbon.		Manganous oxide.		Carbonic oxide.

(2.)	MnO	+	SO_2Ho_2	=	SO_2Mno''	+	OH_2.
	Manganous oxide.		Sulphuric acid.		Manganous sulphate.		Water.

The solution is evaporated to dryness, and again exposed to a red heat. Water is then added, which dissolves out the pure manganous sulphate. Or (2) by heating manganic oxide with sulphuric acid. (See page 66.) The salt was formerly used to produce a permanent brown dye. It forms, like magnesic sulphate, a double salt with potassic sulphate. It crystallizes with 5 molecules of water.

Manganous Carbonate (manganese spar), $COMno''$.—This compound occurs in nature, generally accompanying *spathic iron ore*. It may be obtained in the hydrated form as a white precipitate by adding a solution of sodic carbonate to a solution of manganous sulphate:—

SO_2Mno''	+	$CONao_2$	=	$COMno''$	+	SO_2Nao_2.
Manganous sulphate.		Sodic carbonate.		Manganous carbonate.		Sodic sulphate.

Manganic Chloride, $MnCl_4$.—(See Preparation of Chlorine, page 59.)

Manganic Oxide (black, bin-, deut-, hyper-, or per-oxide of manganese), MnO_2.—This oxide is the most important manganese ore, forming *pyrolusite* when anhydrous, and *wad* when hydrated. It occurs both crystalline and amorphous. It is insoluble in water, and is of a black colour. It may be prepared by adding a solution of sodic or calcic chloro-hypochlorite to a solution of a manganous salt: as thus prepared it is hydrated. Heated with hydrochloric acid, chlorine is evolved, and manganous chloride remains (see p. 59); boiled with sulphuric acid, oxygen is evolved (see p. 66). It is largely employed for preparing chlorine, used in the manufacture of calcic chloro-hypochlorite (bleaching powder).

Dimanganic Trioxide (sesquioxide of manganese), $\left\{\begin{matrix}\mathbf{Mn}\mathrm{O} \\ \mathbf{Mn}\mathrm{O}\end{matrix}\right.\mathrm{O}$.—This compound forms the mineral *braunite;* in the hydrated form it constitutes *manganite* ($\left\{\begin{matrix}\mathbf{Mn}\mathrm{O} \\ \mathbf{Mn}\mathrm{O}\end{matrix}\right.\mathrm{O}, \mathrm{OH}_2$). It may be obtained as a brown hydrate by passing chlorine through water with manganous carbonate suspended in it, and then removing the excess of carbonate by dilute nitric acid. When heated, oxygen is evolved and trimanganic tetroxide remains:—

$$6\left\{\begin{matrix}\mathbf{Mn}\mathrm{O} \\ \mathbf{Mn}\mathrm{O}\end{matrix}\right.\mathrm{O} \quad = \quad 4\left\{\begin{matrix}\mathbf{Mn}\mathrm{O} & \mathrm{O} \\ \mathbf{Mn} & \mathrm{O} \\ \mathbf{Mn}\mathrm{O} & \end{matrix}\right. \quad + \quad \mathrm{O}_2.$$

Dimanganic trioxide. Trimanganic tetroxide. Oxygen.

An oxide known as *varvicite*, and having the comporecently discovered by Mr. W. Weldon,* is prepared by adding an excess of calcic oxide to a solution ot manganous chloride, and forcing air through the mixture :—

$$\underset{\substack{\text{Manganous}\\ \text{chloride.}}}{2\mathbf{Mn}Cl_2} + \underset{\substack{\text{Calcic}\\ \text{oxide.}}}{4\mathbf{Ca}O} + \underset{\substack{\text{Oxygen}\\ \text{(from air).}}}{O_2} = \underset{\substack{\text{Calcic}\\ \text{manganite.}}}{2\mathbf{Mn}OCao''} + \underset{\substack{\text{Calcic}\\ \text{chloride.}}}{2\mathbf{Ca}Cl_2}.$$

The compound is prepared on the large scale from the residual manganous chloride of the chlorine manufactory, and is replacing manganic oxide in that important industry, owing to the following facts, as stated by the discoverer, that "besides being continually reproduced," it "reduces, by fully 80 per cent., the principal item (*i. e.*, the manganic oxide) in the cost of the manufacture of chlorine, greatly increases the quantity of chlorine which can be practically obtained from a given quantity of hydrochloric acid, and, moreover, enables the manufacture of chlorine to be carried on without the production of an offensive residue."

The solution of calcic chloride is run off from the manganite, which is then treated with hydrochloric acid, when chlorine is evolved by the following reaction :—

$$\mathbf{Mn}OCao'' + 6HCl = \mathbf{Mn}Cl_2 + \mathbf{Ca}Cl_2 + 3\mathbf{O}H_2 + Cl_2.$$

The manganous chloride is reconverted into calcic manganite as above, and so the regeneration of the manganite is continuous.

Manganic Acid, $\mathbf{Mn}O_2Ho_2$.—This acid may be obtained by adding dilute sulphuric acid to baric manganate. It is a green liquid, which undergoes spontaneous decomposition very speedily :—

$$\underset{\substack{\text{Baric}\\ \text{manganate.}}}{\mathbf{Mn}O_2Bao''} + \underset{\substack{\text{Sulphuric}\\ \text{acid.}}}{\mathbf{S}O_2Ho_2} = \underset{\substack{\text{Manganic}\\ \text{acid.}}}{\mathbf{Mn}O_2Ho_2} + \underset{\substack{\text{Baric}\\ \text{sulphate.}}}{\mathbf{S}O_2Bao''}.$$

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* *Chemical News*, vol. xx., Sept. 3rd, 1869, p. 109.

Potassic Manganate (manganate of potash), $\mathbf{Mn}O_2Ko_2$.—This is the most important salt of manganic acid. It is prepared by heating to low redness an intimate mixture of 4 parts of powdered manganic oxide, 5 parts of potassic hydrate, and 3½ parts of potassic chlorate; the potassic hydrate is dissolved in the least possible quantity of boiling water and added to the other ingredients, and the mixture gently dried before further heating: the source of heat should be removed from the crucible when the mass has become green all through, which will usually take about an hour. The product should be removed from the crucible when cold, dissolved in water, allowed to stand that the insoluble impurities may deposit, and rendered perfectly clear by filtration—not through paper, however, as that would decompose the salt, but through a funnel, the neck of which is plugged with asbestos. The solution, which is green, when allowed to evaporate spontaneously in a vacuum, yields green crystals having the same form as those of potassic sulphate.

Permanganic Acid (hypermanganic acid), $\left\{\begin{matrix}\mathbf{Mn}O_2(OHo)\\ \mathbf{Mn}O_2(OHo)\end{matrix}\right.$.—If the solution of potassic manganate is boiled or diluted with water, or if a small quantity of acid be added to it, it changes from green to purple, owing to the formation of potassic permanganate$\left(\left\{\begin{matrix}\mathbf{Mn}O_2(OKo)\\ \mathbf{Mn}O_2(OKo)\end{matrix}\right.\right)$ and manganic oxide, or a manganous salt:—

$$\underset{\text{Potassic manganate.}}{3\mathbf{Mn}O_2Ko_2} + \underset{\text{Water.}}{3OH_2} = \underset{\text{Potassic permanganate.}}{\left\{\begin{matrix}\mathbf{Mn}O_2(OKo)\\ \mathbf{Mn}O_2(OKo)\end{matrix}\right\}} +$$

$$\mathbf{Mn}O_2,OH_2 \quad + \quad 4KHO.$$

Manganic oxide-hydrate.	Potassic hydrate.

The acid can be isolated by treating potassic permanganate successively with argentic nitrate, baric chloride, and sulphuric acid. It is a reddish crystalline solid, and its solution when heated to 38° C. evolves oxygen.

Potassic Permanganate, $\left\{\begin{matrix}\mathbf{Mn}O_2(OKo)\\ \mathbf{Mn}O_2(OKo)\end{matrix}\right.$. — This, which is the most important salt of permanganic acid, is prepared from the manganate as already described. Its solution may be boiled without undergoing decomposition, but organic matter decomposes it. If the solid salt is heated, oxygen is evolved, and potassic manganate and manganic oxide left :—

$$2\left\{\begin{matrix}\mathbf{Mn}O_2(OKo)\\ \mathbf{Mn}O_2(OKo)\end{matrix}\right\} = 2\mathbf{Mn}O_2Ko_2 + 2\mathbf{Mn}O_2 + 2O_2.$$

Potassic permanganate.	Potassic manganate.	Manganic oxide.	Oxygen.

Both this salt and the manganate, owing to their great oxidizing power, are used as disinfecting agents, and as oxidizing agents in the laboratory. The permanganate is known as *Condy's fluid*, from Mr. Condy, of Battersea, an extensive maker of it. It is much used in volumetric analysis. The manganate, owing to the change of colour which its solution undergoes on boiling or dilution, is called *mineral chameleon.*

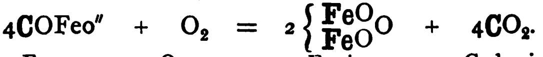
Tests for Manganese Compounds.—Only *manganous* salts are met with in ordinary analysis. Ammonic sulphide precipitates the flesh-coloured manganous sulphide ($\mathbf{Mn}S$), which is soluble in strong acetic acid. Potassic and sodic hydrates precipitate manganous hydrate ($\mathbf{Mn}Ho_2$), which speedily darkens in colour when shaken up in a flask. Fused on platinum-foil with sodic carbonate and a little potassic nitrate, they yield green sodic manganate ($\mathbf{Mn}O_2Nao_2$); this is a characteristic and delicate test. In the oxidizing blowpipe flame, with a borax bead, an amethyst tinge is produced, which colour disappears in the reducing flame.

Iron—Fe.

Atomic weight, 56. Atomicity, rarely VI., most frequently IV. (as a pseudo-triad) and II. Sp. gr., 7·8.

Occurrence in Nature.—Iron occurs to greater or less extent in every country of the world. It occurs in minerals, vegetables, and animals. It is occasionally found in a free state in the mineral kingdom, and also in aërolites or meteoric stones. Its principal and most important ores are *magnetic iron ore*, or *loadstone* ($\left\{\begin{matrix}FeO\\FeO\end{matrix}\right.Feo''$); *specular iron ore*, *fer oligiste*, *red hæmatite*, *micaceous iron*, or *iron glance*, all of which have the same composition, viz., ($\left\{\begin{matrix}FeO\\FeO\end{matrix}\right.O$); *brown hæmatite*, or *fibrous brown iron* cold air was employed, but now the air is heated by passing it through hot pipes: the air by this means is raised to a temperature of 300° or 400° C.; this is called the *hot blast*. This effects a great saving of fuel, and coal may be used instead of coke. The limestone introduced into the furnace separates the aluminic silicate (clay) from the roasted ore, so that the coal can act upon the metallic oxide. The clay combining with the calcic oxide forms a glassy compound of aluminic and calcic silicates, which is called *slag*. The oxide of iron parts with its oxygen to the carbon of the coal. As the iron melts it falls to the bottom of the furnace, uniting in its descent with the white-hot carbon of the coal, forming a carbide of iron, which accumulates in a stratum at the bottom. The slag, having a lower specific gravity than the melted iron, floats like cream upon the surface. The iron and slag are drawn off at different openings at the bottom of the furnace. The iron is run into grooves made of sand, and this constitutes the *pig iron* of commerce. This iron is associated with carbon, and generally with silicon, manganese, magnesium, calcium, sulphur, and phosphorus.

The chemical changes which occur may be explained as follows:—The roasting of the ore converts the ferrous carbonate—for it is of the reduction of this ore, as it is the most common, that we are now speaking—into carbonic anhydride, which escapes as gas, and ferric oxide, which remains:—



$$\underset{\substack{\text{Ferrous}\\\text{carbonate.}}}{4\mathrm{COFeo}''} \quad + \quad \underset{\substack{\text{Oxygen}\\\text{(from air).}}}{\mathrm{O_2}} \quad = \quad \underset{\substack{\text{Ferric}\\\text{oxide.}}}{2\left\{\begin{matrix}\mathrm{FeO}\\\mathrm{FeO}\end{matrix}\right.\mathrm{O}} \quad + \quad \underset{\substack{\text{Carbonic}\\\text{anhydride.}}}{4\mathrm{CO_2}}.$$

The limestone (calcic carbonate) introduced along with the roasted ore into the furnace is speedily [illegible] into [illegible] oxide by the high temperature:—

$$CaCO_3 = CaO + CO_2$$

[illegible] [illegible] Carbonic anhydride.

The [illegible] silica to [illegible] slag, and the ferric oxide is reduced [illegible] by the carbonic oxide [illegible] of the coke:—

$$Fe_2O_3 + 3CO = 2Fe + 3CO_2$$

[illegible] [illegible] Iron. Carbonic anhydride.

[illegible] the iron melts and unites with [illegible] carbon, forming cast iron.

[illegible] of malleable iron [illegible] portion of the carbon it [illegible] by the process called [illegible] the cast iron is laid on [illegible]. The stream of [illegible] the iron, causing it [illegible] by means of a large puddling-rod into large balls, which are hammered by steam hammers. All the liquid slag contained in the balls is squeezed out, and the particles of iron welded together in one mass. This mass is heated a second time and passed between large iron rollers. The fibrous structure of iron is mainly due to this process. Iron which is to sustain great tension is submitted, with slight modification, to this latter process several times. The iron is no longer crystalline ; it is tough and fibrous in its texture, which may be seen by acting on a piece of wrought iron with dilute acid. By violent concussion this toughness and fibrous character disappears, and it again becomes crystalline and brittle. Malleable iron, although much purer than cast iron, is never free from carbon and other substances.

Iron containing a smaller per-centage of carbon than cast iron and a larger per-centage than wrought iron forms the important variety of iron known as *steel*. The purest malleable iron is made into steel by placing bars of iron alternately with charcoal powder in a large brickwork chamber, and exposing the whole to a high red heat for seven or eight days. The iron combines with a small portion of the carbon. It is now more fusible than malleable, but not so fusible as cast iron, and has lost in a great degree its malleability. If fine steel is required, several bars are welded together under the steam hammer, which gives greater uniformity of composition. This process of making steel, however, is to some extent superseded by a method called, from its inventor, the *Bessemer process*. By it steel can be prepared directly from Swedish pig iron. The cast iron is melted and run into a conical iron vessel working on trunnions and open at the top, called a *converter* ; air, at a pressure of about 20 lbs. on the square inch, is then caused [illegible] metal; this causes the [illegible] is the temperature [illegible] that the metal boils. [illegible] of the con- [illegible] brilliant coruscations. [illegible] according to the amount [illegible] a charge of six [illegible] the converter is [illegible] sufficient cast iron [illegible] steel, which is then [illegible] Bessemer cast steel. [illegible] mentioned are pure. [illegible] together in a [illegible] filings and [illegible] ferric oxide in a [illegible] the latter pro- [illegible] the ingots [illegible]

[illegible] is crystalline. [illegible] than wrought iron. [illegible] commerce; [illegible] The [illegible] variety is with nitric acid, while iron gives a green one. Steel is also more brittle and more fusible than iron. Steel can be *tempered*,—that is, its degrees of hardness, elasticity, and brittleness, may be altered by exposing it to varying degrees of heat, and then cooling it either quickly or slowly. *Pure iron.*—Pure iron crystallizes in tetrahedra belonging to the cubic system. With the exceptions of cobalt and nickel, it exceeds in tenacity all other metals; a wire $\frac{1}{86}$ of an inch in diameter will sustain a weight of sixty lbs. It is also very ductile, and is malleable to some extent. Iron does not oxidize or rust in dry air at the common temperature; heated to redness it becomes covered with the black oxide $\left(\left\{\begin{matrix}\mathrm{FeO}\\\mathrm{FeO}\end{matrix}\right.\mathrm{Feo}''\right)$; the same compound is produced when iron wire is burnt in oxygen. Iron at a red heat decomposes water, like potassium at the common temperature, liberating hydrogen, and producing as in the previous cases the black oxide (see p. 53). Dilute sulphuric and hydrochloric acids dissolve iron freely; hydrogen is evolved, and a ferrous salt remains in solution. Iron is converted almost instantly into a magnet when placed in contact with one; it only retains its magnetic character so long as it remains in contact with the magnet; the magnetic iron ore, as its name indicates, is naturally magnetic, its magnetism having been produced by the long-continued inductive influence of the earth's magnetism. Steel does not become magnetic so quickly as iron; but when once it has become so it retains its magnetic power permanently.

Uses.—The principal uses of iron, in its various forms, are chiefly mechanical. It is used chemically, however, in some cases for evolving hydrogen from acids, and as a reducing agent.

Principal Compounds.—The changes which may occur in the active atomicity of iron cause it to form a great many compounds, some of which are of great importance.

Ferrous Compounds (Compounds of Diad Iron).

Ferrous Oxide (protoxide of iron), **FeO**.—This oxide is scarcely known in the free state, owing to its proneness to pass into ferric oxide.

Ferrous Hydrate (hydrated protoxide of iron, hydrated ferrous oxide), **$FeHo_2$**.—If a solution of potassic hydrate be added to a solution of a ferrous salt, out of contact with air, a white precipitate of ferrous hydrate is obtained:—

$$\begin{array}{ccccccc} SO_2Feo'' & + & 2KHo & = & SO_2Ko_2 & + & FeHo_2 \\ \text{Ferrous} & & \text{Potassic} & & \text{Potassic} & & \text{Ferrous} \\ \text{sulphate.} & & \text{hydrate.} & & \text{sulphate.} & & \text{hydrate.} \end{array}$$

This precipitate when exposed to the air becomes first green and then reddish brown, owing to its conversion into ferric oxide by combination with oxygen from the air and elimination of water:—

$$4FeHo_2 + O_2 = 2\left\{\begin{matrix}FeO\\ \end{matrix}\right.O + 4OH_2.$$

By the first process it is obtained in green crystals, having the composition $\mathbf{Fe}Cl_2,4\mathbf{O}H_2$; and by the second it is obtained as a white anhydrous sublimate. It is soluble in alcohol. Heated in air it evolves chlorine and leaves a residue of ferric oxide.

Ferrous Iodide (protiodide of iron), $\mathbf{Fe}I_2$.—This salt is easily prepared by digesting iodine with water and iron:—

$$Fe + I_2 = \mathbf{Fe}I_2.$$

On the solution evaporating spontaneously under the air-pump, it yields green crystals of $\mathbf{Fe}I_2,4\mathbf{O}H_2$. It is used in medicine. Exposed to the air it combines with oxygen; it is therefore mixed with strong syrup, which retards this change.

Ferrous Sulphide (protosulphide or sulphuret of iron), $\mathbf{Fe}S$.—This substance is prepared by throwing into a red-hot crucible, in small quantities at a time, a mixture of two parts of sulphur and three and a half parts of iron filings. It is decomposed by dilute sulphuric and hydrochloric acids, sulphuretted hydrogen being evolved, while ferrous sulphate or chloride remains in solution. It is largely used in the laboratory in preparing sulphuretted hydrogen. If heated in the air it is converted into ferrous sulphate by combination with oxygen from the air.

Ferrous Sulphate (sulphate or protosulphate of iron, copperas, green vitriol), $\mathbf{S}O_2Feo''$; crystallized, $\mathbf{S}OHo_2Feo'',6\mathbf{O}H_2$.—This, which is the commonest and most important ferrous salt, is prepared (1) by the solution of iron in dilute sulphuric acid; (2) by dissolving ferrous sulphide in dilute sulphuric acid; and (3), on the large scale, by the slow oxidation of ferric disulphide ($\mathbf{Fe}S_2$). The disulphide (iron pyrites) is exposed along with scrap iron, to the influence of air [illegible] piece of ground: if sufficient [illegible] does not exist, the bed [illegible] ferrous sulphate and [illegible] the latter reacts upon the [illegible] forming a further quantity [illegible] dissolves the salt, the solution [illegible] evaporated, and on cooling yields [illegible] —

$$[illegible] = 2SO_2Feo'' + 2SO_2Ho_2$$

[illegible] Oxygen. Ferrous sulphate. Sulphuric acid.

$$[illegible] = SO_2Feo'' + H_2$$

[illegible] Hydrogen.

[illegible] blacks: writing ink [illegible] an infusion of galls [illegible] first loses a great [illegible] crystallization, and on continuing [illegible] Nordhausen sulphuric acid [illegible] ferric oxide being left in [illegible]

Ferrous Carbonate (protocarbonate of iron)

Ferric Compounds (Compounds of Tetrad Iron).

Ferric Oxide (peroxide, red oxide, or sesquioxide of iron), $\left\{\begin{matrix}\text{FeO}\\ \text{FeO}\end{matrix}\right.\text{O}$, or $\text{Fe}'''_2\text{O}_3$.—This oxide occurs in nature both crystallized and amorphous, forming the ores of iron known as *red hæmatite*, *specular iron ore*, and *bloodstone*. The crystals are rhombohedral. It is prepared by decomposing ferrous sulphate by heat or by igniting ferric hydrate in a crucible:—

$$\underset{\text{Ferric hydrate.}}{\left\{\begin{matrix}\text{FeHo}_3\\ \text{FeHo}_3\end{matrix}\right.} = \underset{\text{Ferric oxide.}}{\left\{\begin{matrix}\text{FeO}\\ \text{FeO}\end{matrix}\right.\text{O}} + \underset{\text{Water.}}{3\text{OH}_2}.$$

It is a red powder, almost insoluble in acids, and when heated in a current of hydrogen or with carbon is reduced to the metallic state:—

$$1.\quad \underset{\text{Ferric oxide.}}{\left\{\begin{matrix}\text{FeO}\\ \text{FeO}\end{matrix}\right.\text{O}} + \underset{\text{Hydrogen.}}{3\text{H}_2} = \underset{\text{Iron.}}{2\text{Fe}} + \underset{\text{Water.}}{3\text{OH}_2}.$$

$$2.\quad \left\{\begin{matrix}\text{FeO}\\ \text{FeO}\end{matrix}\right.\text{O} + \underset{\text{Carbon.}}{3\text{C}} = 2\text{Fe} + \underset{\text{Carbonic oxide.}}{3\text{CO}}.$$

Under the names of *colcothar*, *crocus of Mars*, and *rouge*, it is used for polishing glass and jewellery; ground with oil it forms a cheap red paint.

Ferric Hydrate (hydrate of sesquioxide as iron), $\left\{\begin{matrix}\text{FeHO}_3\\ \text{FeHO}_3\end{matrix}\right.$.—This compound is prepared by adding ammonic hydrate in excess to a solution of ferric chloride, or any other soluble ferric salt:—

[illegible]

$$[illegible]$$

[illegible]

[illegible]

$$[illegible]$$

[illegible]

[illegible]

oxide is the iron ore known as *loadstone* or *magnetic iron ore*. It is of a lustrous black colour. It is produced when iron is burnt in air or oxygen, and when steam is passed over the red-hot metal. The iron yielded by it is of excellent quality.

Ferric Acid, FeO_2Ho_2.—When one part of ferric oxide and four parts of potassic nitrate are mixed together and heated to redness, a brown substance is obtained, which, when dissolved in water, gives a purple solution of potassic ferrate (FeO_2Ko_2) Potassic ferrate is decomposed by organic matter, and also by acids, with formation of ferric oxide. Ferric acid has not been isolated.

Ferric Disulphide (bisulphide of iron), **FeS_2.**—This occurs as a natural production, forming *iron pyrites*, *martial pyrites*, or *brass lumps*. It crystallizes in cubes, octahedra, and dodecahedra, which are of a brassy lustre. It is formed in some cases by the gradual deoxidation of sulphates by organic matter in the presence of solutions of iron salts. When exposed to the air it oxidizes, forming ferrous sulphate, especially if moistened with water at the same time. It is largely used under the name of *mundic* in the manufacture of sulphuric acid.

Ferric Chloride (sesquichloride, perchloride, or permuriate of iron) $\left\{\begin{matrix}FeCl_3\\FeCl_3\end{matrix}\right.$.—If chlorine in excess be passed over red hot iron filings, or if ferric oxide be dissolved in hydrochloric acid, ferric chloride is produced:—

$$\begin{array}{ccccc}(1)\ 2Fe & + & 3Cl_2 & = & \left\{\begin{matrix}FeCl_3\\FeCl_3\end{matrix}\right. . \\ \text{Iron.} & & \text{Chlorine.} & & \text{Ferric chloride.}\end{array}$$

$$(2)\ \left\{\begin{matrix}FeO\\FeO\end{matrix}\right.O + 6HCl = \left\{\begin{matrix}FeCl_3\\FeCl_3\end{matrix}\right. + 3OH_2.$$

Ferric oxide. Hydrochloric acid. Ferric chloride. Water.

The first process gives it in anhydrous brown scales; the second in solution, which on evaporation yields red crystals of $\left\{\begin{matrix}FeCl_3\\FeCl_3\end{matrix}\right.,6OH_2$; both forms are deliquescent. It is soluble in alcohol: this solution is used in medicine. It is also obtained by oxidizing a solution of ferrous chloride with nitric acid, chlorine, or any other oxidizing agent.

Ferric Sulphate (per-or sesqui-sulphate of iron), $S_3O_6(Fe_2O_6)^{vi}$.—This salt is prepared by boiling a solution of ferrous sulphate acidulated with sulphuric acid, with nitric acid, until no more red fumes are given off; or by dissolving ferric hydrate in sulphuric acid:—

$$(1)\ 2SO_2Feo'' + SO_2Ho_2 + NO_2Ho = S_3O_6(Fe_2O_6)^{vi}$$

Ferrous sulphate. Sulphuric acid. Nitric acid. Ferric sulphate.

$$+ OH_2 + NOHo.$$

Water. Nitrous acid.

$$(2)\ \left\{\begin{matrix}FeHo_3\\FeHo_3\end{matrix}\right. + 3SO_2Ho_2 = S_3O_6(Fe_2O_6)^{vi} + 6OH_2.$$

Ferric hydrate. Sulphuric acid. Ferric sulphate.

It forms a yellowish white deliquescent mass on evaporation, having the composition $S_3O_6(Fe_2O_6)^{vi}$, $9OH_2$, which loses its water when moderately heated. On addition of potassic sulphate to a solution of it, and spontaneous evaporation of the mixture, *potassium-iron alum* is obtained:—

$$S_3O_6(Fe_2O_6)^{vi} + SO_2Ko_2 = S_4O_8(Fe_2O_6)^{vi}Ko_2.$$

Ferric sulphate. Potassic sulphate. Dipotassic-diferric tetrasulphate.

Ferric sulphate is used in dyeing blacks, and is commonly, though *incorrectly*, called *nitrate of iron*, as it is made by the first process given above, in which nitric acid is employed.

Ferric Phosphate, $P_2O_2(Fe_2O_6)^{vi},4OH_2$.—This salt is obtained as a white precipitate on mixing solutions of ferric chloride and hydric-disodic phosphate:—

$$\left\{\begin{array}{l}FeCl_3\\FeCl_3\end{array}\right. + 2PONao_2Ho = 2PO_2(Fe_2O_6)^{vi} + 4NaCl + 2HCl$$

Ferric chloride. Hydric-disodic phosphate. Ferric phosphate. Sodic chloride. Hydrochloric acid.

It is insoluble in acetic acid, but soluble in ferric acetate, and is sometimes produced in estimating phosphoric acid.

Tests for Ferrous Compounds.—Ferrous salts, when hydrated, and their solutions, are of a pale green colour; the anhydrous salts are white.

Sulphuretted hydrogen gives no precipitate in their acidified solutions. Ammonic sulphide throws down the black hydrated sulphide (FeS). Potassic or sodic hydrate precipitates the white ferrous hydrate ($FeHo_2$), which changes to green, and then to brown, on exposure to the air. Potassic ferrocyanide causes a bluish white precipitate of dipotassic-ferrous ferrocyanide ($Fe''Fe''Cy_6K_2$), which becomes of a darker blue by exposure to the air; the precipitate is decomposed with precipitation of ferrous hydrate by potassic hydrate: it is insoluble in dilute acids. Potassic ferricyanide causes a dark blue precipitate of ferrous ferricyanide ($Fe_3Fe_2Cy_{12}$), known as *Turnbull's blue;* it is decomposed, like dipotassic-ferrous ferrocyanide, by potassic hydrate, and is insoluble in acids.

Tests for Ferric Compounds.—Ferric salts in solution have a yellow or reddish brown colour.

Sulphuretted hydrogen reduces ferric salts to ferrous salts, with a separation of sulphur, which renders the liquid milky:—

$$2\left\{\begin{matrix}\mathbf{Fe}Cl_3\\ \mathbf{Fe}Cl_3\end{matrix}\right. + 2\mathbf{S}H_2 = 4\mathbf{Fe}Cl_2 + 4HCl + S_2.$$

Ferric chloride. Sulphuretted hydrogen. Ferrous chloride. Hydrochloric acid. Sulphur.

Ammonic sulphide precipitates ferrous sulphide, as ferric sulphide cannot be formed in this way:—

$$2\left\{\begin{matrix}\mathbf{Fe}Cl_3\\ \mathbf{Fe}Cl_3\end{matrix}\right. + 6\mathbf{S}Am_2 = 4\mathbf{Fe}S + S_2 + 12AmCl.$$

Ferric chloride. Ammonic sulphide. Ferrous sulphide. Sulphur. Ammonic chloride.

Soluble hydrates precipitate brown ferric hydrate $\left(\left\{\begin{matrix}\mathbf{Fe}Ho_3\\ \mathbf{Fe}Ho_3\end{matrix}\right.\right)$; insoluble in excess of the reagent, soluble in acids. Potassic ferrocyanide causes a dark blue precipitate of ferric ferrocyanide $(\mathbf{Fe}^{iv}_2[FeCy_6]^{iv}_2)$,—commonly called *Prussian blue*, insoluble in acids, decomposed by soluble hydrates. Potassic ferricyanide produces no precipitate, but the solution becomes deeper in colour. Potassic sulphocyanide changes the colour of ferric salts to a deep blood-red, owing to the formation of a soluble ferric sulphocyanide. This is a delicate reaction.

Cobalt—Co.

Atomic weight, 58·8. Atomicity, VI., (?), IV., and II. Sp. gr., 8·5.

Special Notice.—Cobalt was discovered in 1733, by Brandt.

Occurrence in Nature.—This metal occurs in the metallic state in meteoric stones along with nickel, iron, &c. Its principal ores are *speiss cobalt*, or *tin-white cobalt*, $\left\{\begin{matrix}''\mathrm{As}'\\ ''\mathrm{As}\end{matrix}\right.\mathrm{Co}''$; as *cobalt glance*, and *cobalt pyrites* $\left(\left\{\begin{matrix}\mathrm{CoS}_{\mathrm{S}}\\ \mathrm{CoS}^{\mathrm{S}}\end{matrix}\right.\right)$.

Preparation.—The ore is roasted to expel the arsenic; dissolved in *aqua regia*, evaporated to dryness, re-dissolved in water, and sulphuretted hydrogen passed through the liquid to precipitate copper, bismuth, and arsenic. The filtrate is then boiled with nitric acid to convert the ferrous salts present into ferric salts; ammonic hydrate is then added, which precipitates ferric hydrate, while the cobalt, together with any nickel which the ore may have contained, remains in solution. These two metals are then precipitated together as sulphides, the sulphides re-dissolved in nitric acid, and treated, after dilution, with a current of chlorine gas, or a solution of bromine, and afterwards digested for twelve hours on baric carbonate; the cobalt is gradually precipitated as cobaltic oxide $\left(\left\{\begin{matrix}\mathrm{CoO}_{\mathrm{O}}\\ \mathrm{CoO}^{\mathrm{O}}\end{matrix}\right.\right)$. The mixture of excess of baric carbonate and cobaltic oxide is then dissolved in hydrochloric acid, the barium precipitated as sulphate by adding sodic sulphate, and the cobalt precipitated as cobaltous hydrate ($\mathrm{CoHo_2}$) by sodic hydrate, which is then reduced in a current of hydrogen. The process is very complicated.

Properties.—Cobalt is a hard, infusible, strongly magnetic, reddish grey metal. It is soluble in sulphuric and hydrochloric acids, with evolution of hydrogen; nitric acid also dissolves it freely. It oxidizes by exposure to the air. It is more tenacious than iron.

Uses.—Cobalt is not used in the metallic state. Several of its compounds, however, are esteemed as colours.

Principal Compounds.

Cobaltous Compounds (Compounds of Diad Cobalt).

Cobaltous Oxide (protoxide of cobalt), **CoO**.—This compound is obtained by carefully heating cobaltous hydrate:—

$$CoHo_2 = CoO + OH_2$$

It is of a greenish grey colour. It dissolves in acids, forming cobaltous salts, which are blue in concentrated solutions, and pink when dilute. The *zaffre* of commerce is impure cobaltous oxide. It is used in preparing *smalt*, a beautiful blue glass, used in staining paper, and for other purposes.

Cobaltous Hydrate (hydrated oxide of cobalt), **$CoHo_2$**.—This substance is obtained as pale rose-coloured precipitate, on adding potassic hydrate to a solution of a cobaltous salt:—

$$CoCl_2 + 2KHo = CoHo_2 + 2KCl.$$

Cobaltous Nitrate (nitrate of cobalt), $\left\{\begin{matrix}NO_2\\ Coo''\\ NO_2\end{matrix}\right.$— This salt is prepared by dissolving the oxide in nitric acid :—

$$\underset{\text{Cobaltous oxide.}}{CoO} \quad + \quad \underset{\text{Nitric acid.}}{2NO_2Ho} \quad = \quad \underset{\text{Cobaltous nitrate.}}{\left\{\begin{matrix}NO_2\\ Coo''\\ NO_2\end{matrix}\right.} \quad + \quad \underset{\text{Water.}}{OH_2}.$$

It is a deliquescent salt. It is used for detecting compounds of certain metals with which it gives characteristic coloured compounds when heated on charcoal before the blowpipe; for example, magnesium compounds become pink, those of aluminium blue, and zinc compounds become green.

Cobaltic Oxide (sesquioxide of cobalt), $\left\{\begin{matrix}CoO\\ CoO\end{matrix}\right. O$.— This oxide is obtained by suspending cobaltous hydrate in water or potassic hydrate, and passing a current of chlorine through the liquid; cobaltic hydrate is precipitated, and when gently heated is converted into cobaltic oxide.

It forms a black powder, which, when strongly heated, loses oxygen, and forms cobaltous-dicobaltic tetroxide $\left(\left\{\begin{matrix}CoO\\ CoO\end{matrix}\right. Coo''\right)$, corresponding to magnetic oxide of iron.

Tests for Cobalt Compounds.—All the cobaltous salts (which are the only ones of importance) are lilac-coloured when anhydrous; their concentrated solutions are blue, and dilute solutions are pink or red. Sulphuretted hydrogen gives no precipitate in acid solutions. Ammonic sulphide precipitates a black hydrated sulphide (CoS). Potassic hydrate precipitates first a blue basic salt, which an access of the reagent turns into the rose-coloured hydrate $\mathrm{CoHo_2}$). Before the blowpipe, cobalt compounds give in the oxidizing flame a very characteristic blue bead with borax.

Nickel—Ni.

Atomic weight, 58·8. Atomicity, VI. (?), IV., and II.
Sp. gr., 8·7.

Special Notice.—Nickel was discovered by Cronstedt in 1751. It has a remarkable resemblance to cobalt, its atomic weight and atomicity being the same; and it is always associated with it in nature.

Occurrence in Nature.—Nickel occurs in the metallic state in meteoric stones; its principal ore, however, is *Kupfernickel* ($\left\{\begin{matrix}{}''\mathbf{As}'\\ {}''\mathbf{As}'\end{matrix}\right.\mathrm{Ni}''$, nickelous diarsenide).

Preparation.—Nickel is prepared either from *Kupfernickel* or *speiss* (impure nickelous sulphide-diarsenide, $\left\{\begin{matrix}\mathbf{As}\\ \mathbf{As}\end{matrix}\right.\mathrm{S''Ni''}$). The process employed at Birmingham for extracting it from speiss on the large scale, consists in first fusing it with calcic carbonate and calcic fluoride, grinding the resulting mass to powder, and heating it for twelve hours to expel the arsenic. The residue is dissolved in hydrochloric acid, the solution diluted, and the iron present converted into ferric salt by cautiously adding *bleaching powder:* calcic hydrate (solid) is then added to remove the iron as ferric hydrate. The precipitate is removed by filtration, and sulphuretted hydrogen passed through the filtrate; this precipitates copper, lead, and bismuth as sulphides, leaving only nickel and cobalt in solution. After again filtering and boiling to expel excess of sulphuretted hydrogen, calcic hydrate is added till the liquid is neutral; and on now adding bleaching powder, cobaltic hydrate is precipitated, after removing which the nickel is precipitated as nickelous hydrate ($NiHo_2$). The hydrate is mixed with charcoal, and heated in a crucible for an hour, a button of nickel containing carbon remains at the bottom of the crucible.

Properties.—Nickel is a white, lustrous, hard, ductile metal. It is rather more fusible than iron, but exceeds it in tenacity. It is capable of being rendered magnetic. It is slowly soluble in hydro-chloric and dilute sulphuric acids, hydrogen being evolved; nitric acid and *aqua regia* dissolve it rapidly. Heated in air it becomes oxidized. If *brass* be melted with nickel a white alloy known as *German silver* is obtained.

Uses.—Nickel is principally used in the manufacture of *German* or *nickel silver*, the alloy just mentioned—good proportions for making which are, in 100 parts, copper 51 parts, zinc 30·6, and nickel, 18·4.

Principal Compounds.

Nickelous Compounds (Compounds of Diad Nickel).

Nickelous Oxide (protoxide of nickel), **NiO**.—This oxide is prepared by igniting the hydrate, nitrate, or carbonate in a covered crucible.

It is of an olive-green colour, and forms nickelous salts by solution in acids.

Nickelous Hydrate (hydrate of protoxide of nickel), **$NiHo_2$**.—This substance is obtained as a light green bulky powder by adding an excess of potassic hydrate to a soluble nickelous salt, and washing the resulting precipitate well with boiling water. Ammonic hydrate dissolves it, forming a deep blue solution.

Nickelous Sulphide (protosulphide of nickel),

NiS.—This compound is found in nature as *millerite*, *capillary pyrites*, or *hair nickel.*

Nickelous Chloride (chloride of nickel), $NiCl_2$.—This salt is formed by dissolving nickelous oxide or carbonate in hydrochloric acid. A green solution is obtained, which, on evaporation, yields crystals of the same colour, having the composition $NiCl_2,9OH_2$. If these crystals be heated, water escapes, and a yellowish brown mass of the anhydrous chloride remains.

Nickelous Sulphate (sulphate of nickel), SO_2Nio''.—This salt is prepared by dissolving nickel, nickelous oxide, or carbonate, in sulphuric acid :—

$$\begin{array}{ccccccc} Ni & + & SO_2Ho_2 & = & SO_2Nio'' & + & H_2. \\ \text{Nickel.} & & \text{Sulphuric acid.} & & \text{Nickelous sulphate.} & & \text{Hydrogen.} \end{array}$$

It is the most important of the salts of nickel. It crystallizes in green rhombic prisms, whose composition is $SOHo_2Nio'',6OH_2$, and which change into octahedra by exposure to light.

Nickelic Compounds (Compounds of Tetrad Nickel).

Nickelic Oxide (sesquioxide of nickel), $\left\{\begin{matrix} NiO \\ NiO \end{matrix}\right. O$, may be obtained by gently heating nickelic hydrate :—

$$\begin{array}{ccccc} \left\{\begin{matrix} NiHo_3 \\ NiHo_3 \end{matrix}\right. & = & \left\{\begin{matrix} NiO \\ NiO \end{matrix}\right. O & + & 3OH_2. \\ \text{Nickelic hydrate.} & & \text{Nickelic oxide.} & & \text{Water.} \end{array}$$

It forms a black powder, which is decomposed by a strong heat or by acids, giving off oxygen, and forming nickelous salts with the latter.

Nickelic Hydrate (hydrated sesquioxide of nickel,) $\left\{\begin{matrix} NiHo_3 \\ NiHo_3 \end{matrix}\right.$, is prepared by treating nickelous hydrate suspended in water with a solution of *bleaching powder* :—

$3\mathbf{Ni}Ho_2$	+	Cl_2	=	$\left\{\begin{array}{l}\mathbf{Ni}Ho_3\\ \mathbf{Ni}Ho_3\end{array}\right.$	+	$\mathbf{Ni}Cl_2.$
Nickelous hydrate.		Chlorine.		Nickelic hydrate.		Nickelous chloride.

It forms a black powder, which is decomposed by acids similarly to the oxide.

Tests for Nickel Compounds.—Anhydrous nickelous salts are yellowish brown; hydrated ones green.

Sulphuretted hydrogen gives no precipitate in acidified solutions. Ammonic sulphide precipitates the black hydrated sulphide ($\mathbf{Ni}S$), slightly soluble in excess of the reagent. Potassic hydrate in excess throws down nickelous hydrate ($\mathbf{Ni}Ho_2$) of a pale green colour. Before the blowpipe, when fused with *borax* or *microcosmic salt*, nickel compounds give a reddish yellow glass in the outer flame; the colour fades on cooling. In the inner flame greyish particles of metallic nickel are produced in the bead.

EXERCISES.

THE equation explaining the reaction will furnish the data for the calculations.

The atomic weights on pp. 3 and 4 must be used.

The molecular weight of a compound in criths, with rare exceptions, measures two litres at the normal temperature and pressure (0° C. and 760 m.m. Bar.).

SERIES I.

*1. Classify the following substances into elements and compounds:—water, chalk, common salt, iron, mercury, sulphur, calomel, copperas, ammonia, oxygen, lime, sand.

*2. Why is hydrogen regarded as an element, and hydrochloric acid as a compound?

3. Classify the following elements (1) into chlorous or negative, and basylous or positive; (2) into metals and non-metals:—

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* Science and Art Department Examination Paper, 1869.

iodine, sulphur, carbon, selenium, sodium, barium, fluorine, zinc, quicksilver, hydrogen, and phosphorus.

4. Explain the fact that zinc liberates tin from its solutions, iron liberates copper, and chlorine displaces iodine.

5. Why does mercuric oxide break up into oxygen and mercury when strongly heated?

SERIES II.

6. Define the terms "atom" and "molecule" in their modern acceptation.

7. What is the weight of a litre of oxygen, nitrogen, iodine vapour, and chlorine, respectively, each being measured at the normal temperature and pressure?

8. Give examples of compound molecules which contain two, three, four, five, and six volumes of elementary bodies respectively. How many volumes does a molecule measure when in the gaseous state?

9. If a cubic foot of steam has all the oxygen removed from it, how much will the remaining hydrogen measure?

10. If a cubic foot of ammonia gas has all the nitrogen removed from it, how much will the residual hydrogen measure?

SERIES III.

11. Describe clearly what the following formulæ express:—

$$OH_2,\ NH_3,\ 7OH_2,\ (NH_3)_{10}.$$

12. Platinum forms two combinations with chlorine, PtCl and $PtCl_2$; by what names do you distinguish these two chlorides? Give examples of similar cases.

17. One atom of zinc is found to displace two atoms of hydrogen, and five atoms of calcium exchange places with two atoms of antimony: explain this.
18. Give graphic representations of an atom of phosphorus, silicon, carbon, boron, selenium, mercury, and chlorine, showing how many bonds each atom has. Express the same thing by means of Dr. Odling's atomicity marks.
19. Classify the elements in Exercise 3 according to their atomicities.
20. Explain why phosphorus is frequently a triad as well as a pentad, but never a tetrad or diad.

SERIES V.

21. Define a compound radical; mark the atomicity of the following radicals:—Zno, Cao, Bio, Amo, Am, Auo, Al_2o; give the symbols for the radicals $(SnO_2)''$ and $(SnO_4)^{IV}$.
22. Write out the graphic formulæ of the following compounds:—

$$\mathbf{S}O_2Cuo'', \mathbf{Cr}O_2Ho_2, \mathbf{S}OHo_2, \mathbf{Ba}Ho_2, \left\{\begin{array}{l}\mathbf{N}O_2\\ Bao''\\ \mathbf{N}O_2\end{array}\right., \ '\mathbf{Fe}'''_2Cl_6,$$

$$\mathbf{Cr}_2HO_6, \mathbf{S}_3O_6\mathbf{Fe}_2o^{vi}.$$

23. Write out constitutional formulæ of the following salts:—cadmic bromide, calcic chlorate, aluminic chloride, stannic chloride, strontic nitrate, ferric sulphate, ferrous sulphate, calcic orthophosphate, magnesic silicate.*

Name the following salts:—$\mathbf{S}O_2Cao''$, $\mathbf{S}OCao''$, $\mathbf{S}S''ONao_2$, $\left\{\begin{array}{l}\mathbf{O}Cl\\ \mathbf{O}Ko\end{array}\right.$, $\mathbf{N}ONao$, $Cao''Cl_2$, $\left\{\begin{array}{l}\mathbf{C}OAmo\\ \mathbf{C}OAmo\end{array}\right.$, $\mathbf{Cr}O_2Ko_2$, $\mathbf{C}OCao''$, $\mathbf{As}OKo_3$, $\mathbf{Si}Mno''_2$, $\mathbf{S}_3O_6\mathbf{Al}_2o^{VI}$.

24. What are tribasic acids? Give examples.
25. What is the difference between atomic and molecular combination? Give examples.

SERIES VI.

26. What weight of hydrogen gas will be evolved by the solution of 100 grammes of zinc in dilute sulphuric acid?†

Ans. 3·07 grammes.

27. How many litres will the hydrogen of (26) measure at the normal temperature and pressure? Ans. 34·34 litres.

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* See table of salts, p. 33.

† *Example* 12. What weight of hydrogen will be evolved by acting on hydrochloric acid with 10 grammes of magnesium?

By the equation,—

28. If 10 grammes of sodium (Na) be placed upon water, how much sodic hydrate (NaHo) will be formed?
Ans. 17·39 grammes.

29. If steam be passed through a red-hot porcelain tube containing 10 grammes of iron filings until they become completely converted into magnetic oxide of iron (Fe_3O_4), how many litres of hydrogen will be obtained?
Ans. 5·31 litres.

30. Sixty-five grammes of zinc are dissolved by boiling in potassic hydrate (KHo), and the hydrogen evolved burnt in oxygen: how much water will be produced? Ans. 18 grammes.

SERIES VII.

31. How many litres of chlorine will be evolved by acting on a kilo. of manganic oxide (MnO_2) with hydrochloric acid?
Ans. 256·56 litres.

32. How much argentine chloride (AgCl) will be formed by adding an excess of argentic nitrate (NO_2Ago) to a solution of 10 grammes of common salt (NaCl)?
Ans. 24·53 grammes.

33. How much salt cake (SO_2Nao_2) can be obtained from a ton of salt (NaCl)? Ans. 1·213 tons.

$$\underset{24}{\underset{|}{Mg}} + 2HCl = MgCl_2 + \underset{2}{\underset{|}{H_2}}.$$

$$\begin{array}{lccccccc} & \text{Mg.} & & \text{Mg.} & & \text{H.} & & \text{H.} \\ \therefore \text{ as} & 24 & : & 10 & :: & 2 & : & x. \end{array}$$

$$x = \frac{10 \times 2}{24} = \cdot 83 \text{ gramme of hydrogen.}$$

Example 13. How many litres of hydrogen, at 0° C. and 760 m.m. Bar., will be evolved in (28)?

By the equation,—

$$\underset{46}{\underset{|}{Na_2}} + 2OH_2 = 2NaHo + \underset{2}{\underset{|}{H_2}}.$$

Therefore 46 criths of sodium yield 2 criths of hydrogen.

2 criths of hydrogen = 2 litres.

$$\begin{array}{lccccccc} & \text{Grammes Na.} & & \text{Gr. Na.} & & \text{Lit.} & & \text{Lit.} \\ \therefore \text{ As} & 46 \times \cdot 0866 & : & 10 & :: & 2 & : & x. \end{array}$$

$$x = \frac{10 \times 2}{46 \times \cdot 0896} = 4 \cdot 85 \text{ litres of hydrogen.}$$

34. How much common salt, sulphuric acid, and manganic oxide respectively will it require to produce 100 litres of chlorine gas? Ans. $NaCl$, 524·16 grammes.
SO_2Ho_2, 887·08 „
MnO_2, 389·76 „

35. How much silver is contained in 100 grammes of argentic chloride? Ans. 75·26.

SERIES VIII.

36. A gas bag holding 20 litres is required to be filled with oxygen at the normal temperature and pressure: how much potassic chlorate must be used?
Ans. 73·173 grammes.

Example 14. How much sodium is contained in 10 grammes of common salt?

By the formula,—

$$\underbrace{\overset{\text{Na.}}{23} + \overset{\text{Cl.}}{35{\cdot}5}}_{58{\cdot}5}.$$

$$\therefore \text{ as } \overset{\text{NaCl.}}{58{\cdot}5} : \overset{\text{NaCl.}}{10} :: \overset{\text{Na.}}{23} : x.$$

$$x = \frac{10 \times 23}{58{\cdot}5} = 3{\cdot}93 \text{ grammes of sodium.}$$

See also *Example* 11, p. 122; and on p. 123, for 37·9 litres of HCl, read 38·15 litres.

Example 11 may also be worked as follows:—

By the equation, 117 criths of NaCl yield 73 criths of HCl.

73 criths = 4 litres.

$$\therefore \text{ as } \overset{\text{Grammes NaCl.}}{117 \times {\cdot}0896} : \overset{\text{Gr.NaCl.}}{100} :: \overset{\text{Lit.HCl.}}{4} : \overset{\text{Lit.HCl.}}{x}.$$

$$x = \frac{100 \times 4}{117 \times {\cdot}0896} = 38{\cdot}15 \text{ litres of HCl.}$$

Example 15. If 10 litres of chlorine be mixed with 10 litres of sulphuretted hydrogen, how much hydrochloric acid gas will be produced?

By the equation,—

$$2SH_2 + 2Cl_2 = 4HCl + S_2.$$

2 molecules 2 molecules 4 molecules
(4 volumes). (4 volumes). (8 volumes).

The chlorine forms double its volume of hydrochloric acid.

∴ 10 litres of chlorine will form 20 litres of hydrochloric acid.

37. From a given weight of manganic oxide (MnO_2), how much more oxygen can be obtained by heating it with sulphuric acid than by heating it alone? Ans. $\frac{1}{2}$ more.

38. If 5 litres of chlorine gas be mixed with steam, and the mixture sent through a red-hot porcelain tube, how many litres of oxygen will be liberated? Ans. 2·5 litres.

39. 100 cubic metres of oxygen are required: how much sulphuric acid must be decomposed in order to furnish it (see p. 68)? Ans. 878,080 grammes.

40. 100 c.c. of ozone are heated strongly: how many c.c. of oxygen will be obtained? Ans. 150 c.c.

SERIES IX.

41. If 28 c.c. of hydrogen be mixed with 15 c.c. of oxygen, how many c.c. of water vapour will be formed, calculated at 0° C. and 760 m.m. Bar., and what will it weigh?
Ans. 28 c.c., and ·0225792 gramme.

42. What is the weight of a litre of water vapour, calculated at 0° C. and 760 m.m. pressure? Ans. 9 criths.

43. If the hardness of a water showing 12 degrees of hardness with Clark's soap-test, arise wholly from the presence of calcic sulphate (SO_2Cao'') in solution, how many grains of this salt per gallon does the water contain?
Ans. 16·32 grammes of SO_2Cao''.

*44. How is hydroxyl prepared? what are the relations of this substance to water and slaked lime? and to what useful purpose has it been applied?

45. What weight of pure hydroxyl (Ho_2) can be obtained from 1 lb. of baric peroxide (BaO_2)? Ans. ·2011 lb.

SERIES X.

46. How much bleaching powder ($Ca[OCl]Cl$) can be made with a ton of lime (CaO)? Ans. 2·267 tons.

47. How many litres of chlorine will it require to convert 1 kilogramme of lime (CaO) into bleaching powder?
Ans. 398·59.

48. How many litres of chlorine are require to furnish a kilo. of potassic chlorate? (see pp. 83 and 176).
Ans. 546·64 litres.

49. What are meant by allotropic modifications? Give numerous examples.

50. Write out the constitutional and graphic formulæ of the compounds of chlorine wlth oxygen and hydroxyl.

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* *Science* and Art Department Examination Paper, 1869.

SERIES XI.

51. How much dry calcic chloride ($CaCl_2$) can be obtained by dissolving a kilo. of chalk in hydrochloric acid?
Ans. 1,110 grammes.

52. How many litres will the carbonic anhydride evolved in (51) measure at 0° C. and 760 m.m. Bar.?
Ans. 223·16 litres.

53. If 100 litres of carbonic oxide be mixed with 50 litres of oxygen gas, and the mixture exploded, how many litres of carbonic anhydride will be formed? Ans. 100 litres.

54. If 100 litres of carbonic anhydride be passed through a red-hot tube containing charcoal, how many litres of carbonic oxide will be obtained? Ans. 200 litres.

55. How many litres of carbonic oxide can be obtained from 100 grammes of dry oxalic acid, $\left\{\begin{array}{l}COHo\\COHo\end{array}\right.$?
Ans. 24·8 litres.

SERIES XII.

56. The density of atmospheric air is 14·4: show that this is also the calculated density of dry air, air being considered to be a mixture.

57. How many litres of ammonia gas can be obtained from 100 grammes of sal ammoniac (NH_4Cl)? Ans. 41·72 litres.

58. How much cupric nitrate, $\left\{\begin{array}{l}NO_2\\Cuo''\\NO_2\end{array}\right.$, can be obtained by dissolving 10 grammes of metallic copper in nitric acid (NO_2Ho)? Ans. 29·52 grammes.

59. What is the weight of 10 litres of ammonia gas, calculated by Williamson's standard? Ans. 7·59 grammes.

60. How much more nitric acid (NO_2Ho) can be obtained from a kilo. of sodic nitrate (NO_2Nao) than from a kilo. of potassic nitrate (NO_2Ko)? Ans. 117·41 grammes.

SERIES XIII.

61. How many litres of laughing gas (ON_2) can be obtained from a kilo. of ammonic nitrate (NO_2Amo)?
Ans. 279·01 litres.

62. Ten litres of gas are measured off at 0° C. and 760 m.m. Bar.: how many litres will they measure if raised to 100° C. and under a pressure of 750 m.m.? (see p. 117).
Ans. 13·84 litres.

*63. I measure 103 c.c. of dry nitrogen at 10° C. and 650 m.m. mercurial pressure: what will be its volume at 0° C. and 760 m.m. pressure, and how much will it weigh in grammes? Ans. 82·5 c.c.

64. What is the theoretical quantity of pure nitric acid obtainable from one ton of sodic nitrate? Ans. ·7411 ton.

*65. I take equal volumes (measured at the same temperature and pressure) of hydrogen, nitrogen, hydrochloric acid, gas, steam, and carbonic anhydride, and find that the nitrogen weighs 56 grammes. Required the weight of each of the other gases.
Ans. $H=4$, $HCl=73$, $OH_2=36$, $CO_2=88$ grammes.

SERIES XIV.

66. The iodine in 100 volumes of hydriodic acid is liberated in succession by chlorine and oxygen: how many volumes of chlorine and of oxygen are required?
Ans. 50 volumes of chlorine and 25 vols. of oxygen.

67. 50 litres of gas are measured off at 13° C.: what will they measure at −40° C., the pressure remaining the same?
Ans. 40·73 litres.

68. To what temperature must 18 litres of oxygen be raised in order that it may fill the bag in Exercise 36?
Ans. 30⅓° C.

69. 100 c.c. of gas are measured off at 20° C. and 755 m.m. Bar.: how much will they measure at—20° C. and 765 m.m. Bar? Ans. 85·21 c.c.

70. To what pressure must 80 litres of gas measured at 760 m.m. Bar. be reduced in order that it may measure 100 litres, the temperature remaining the same?
Ans. 608 m.m. Bar.

SERIES XV.

71. What is the weight in criths and in grammes of 1 litre each of nitrous oxide (ON_2), carbonic anhydride (CO_2), carbonic oxide (CO), and iodine vapour?
Ans. Criths $ON_2=22$, $CO_2=22$, $CO=14$, $I=127$,
Grammes $ON_2=1·97$, $CO_2=1·97$, $CO=1·25$,
$I=11·37$.

72. 80 criths of dry ammonic nitrate yield 2 litres of nitrous oxide gas (ON_2): how many litres can be obtained from 100 grammes of ammonic nitrate? Ans. 27·9 litres.

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* Science and Art Department Examination Paper, 1869.

73. How many litres of sulphurous anhydride (SO_2) would be obtained by dissolving 100 grammes of copper in sulphuric acid (SO_2Ho_2)? Ans. 35·15 litres.

74. How much crystallized cupric sulphate ($SO_2Cuo'',4OH_2$) would be produced in (73)? Ans. 364·56 grammes.

75. In a vessel containing 100 litres of air, sulphur is burnt until the whole of the oxygen is consumed: what will then be the composition of the atmosphere by volume?
Ans. $SO_2 = 20$ litres; $N = 80$ litres.

SERIES XVI.

76. If 64 lbs. of sulphur be burnt in air, how many lbs. of sulphurous anhydride (SO_2) will be obtained?
Ans. 128 lbs.

77. What weight of sulphuretted hydrogen (SH_2) will be evolved by acting on 1 kilo. of ferrous sulphide (FeS) with dilute sulphuric acid? Ans. 386·36 grammes.

78. How many litres will the gas evolved in (77) measure at 0° C. and 760 m.m. Bar.? Ans. 253·65 litres.

79. How much sulphuric acid (SO_2Ho_2) can be manufactured from 1 ton of pure iron pyrites (FeS_2), supposing all the sulphur to be available? Ans. 1·63 tons.

80. How much sulphur is contained in 100 grammes of sulphuric acid (SO_2Ho_2)? Ans. 32·65 grammes.

SERIES XVII.

81. 100 grammes of baric sulphate (SO_2Bao'') are precipitated from a solution of Glauber's salt ($SO_2Nao_2,10OH_2$) by means of an excess of a solution of baric chloride ($BaCl_2$): what weight of Glauber's salt was in solution?
Ans. 138·19 grammes.

82. How much cupric sulphide (CuS) will a litre of sulphuretted hydrogen (SH_2) precipitate from an acid solution of cupric chloride ($CuCl_2$)? Ans. 4·2784 grammes.

83. How much sulphur is contained in a kilo. of carbonic disulphide, CS_2? Ans. 842·105 grammes.

84. How many litres of carbonic and sulphurous anhydrides respectively will be produced by the complete combustion in air or oxygen of 100 grammes of carbonic disulphide?
Ans. 29·37 litres CO_2 and 58·74 litres SO_2.

85. How much sulphuric acid (SO_2Ho_2) can be obtained from 1 ton of crude sulphur, containing 90 °/ₒ of pure sulphur?
Ans. 2·75625 tons.

SERIES XVIII.

*86. If 50 grammes of sodium be heated in a current of hydrobromic acid gas until it is completely converted into sodic bromide, what will be the weight and volume of the hydrogen liberated? Ans. 2·173 grammes; 24·26 litres.

87. A current of sulphuretted hydrogen is passed into water containing 80 grammes of bromine, until the bromine is converted into hydrobromic acid: how much sulphur will be obtained? Ans. 16 grammes.

88. How much iodine will be separated from a solution of potassic iodide by passing a litre of chlorine gas into it? Ans. 11·3792 grammes.

89. How much hydrofluoric acid (HF) can be obtained by treating a kilo. of cryolite ($Al_2F_6,6NaF$) with sulphuric acid? Ans. 570 grammes.

90. How much silica (SiO_2) will be required to furnish 100 litres of silicic fluoride (SiF_4)? Ans. 271·04 grammes.

SERIES XIX.

91. How many litres of phosphorous vapour, calculated at 0° C. and 760 m.m. Bar., can be obtained from a kilo. of pure bone ash ($P_2O_2Cao''_3$), supposing all the phosphorus available? Ans. 36 litres.

92. How many litres of phosphoretted hydrogen (PH_3) can be obtained by boiling 10 grammes of phosphorus with sodic hydrate (NaHo)? Ans. 1·8 litres.

93. How much crystallized rhombic sodic phosphate ($PONao_2Ho$,

*97. You have mercury, a glass flask, and a piece of hard glass tube, and are required to make pure oxygen: how will you do it?

*98. What is the weight in grammes of one litre each of the following gases and vapours measured at 0° C. and 760 m.m. Bar.?—nitrogen, arsenic, mercury, mercuric chloride, sulphurous anhydride, ammonia.

Ans. $N = 1{\cdot}25$ grammes.
$As = 13{\cdot}44$,,
$Hg = 8{\cdot}96$,,
$HgCl_2 = 12{\cdot}14$,,
$SO_2 = 2{\cdot}86$,,
$NH_3 = {\cdot}76$ gramme.

99. How much calcic carbonate ($COCao''$) must be decomposed in order to furnish sufficient carbonic anhydride to convert 10 grammes of potassic hydrate (KHo) into potassic carbonate ($COKo_2$)? Ans. 8·928 grammes.

100. How much metallic potassium will be required to furnish 100 grammes of potassic hydrate (KHo)?
Ans. 69·64 grammes.

SERIES XXI.

101. One ton of common salt has to be converted into crystallized sodic carbonate by Le Blanc's process: what weight of $CONao_2, 10OH_2$ will be obtained, supposing there is no loss? Ans. 2·44 tons.

102. Show by an equation what takes place when chlorine gas is passed into a solution of sulphurous acid ($SOHo_2$).

103. One gramme of ammoniochloride of platinum ($2AmCl, PtCl_4$) is ignited: how much metallic platinum will be left? Ans. ·441 gramme.

104. If 100 grammes of metallic zinc be placed in solution of cupric sulphate (SO_2Cuo''), how much metallic copper will separate? Ans. 97·69 grammes.

105. How many grammes of metallic lead can be got out of 100 grammes of red-lead (Pb_3O_4)?
Ans. 90·656 grammes.

SERIES XXII.

106. What are the densities of carbonic oxide (CO), carbonic anhydride (CO_2), and hydrochloric acid gas (HCl), referred to hydrogen as unity?
Ans. $CO = 14$, $CO_2 = 22$, $HCl = 18{\cdot}25$.

Science and Art Department Examination Paper, 1869.

107. If 10 c. c. of ammonia gas be decomposed by the electric discharge, how many c.c. of nitrogen and hydrogen respectively will be obtained? Ans. Nitrogen = 5 c.c.
Hydrogen = 15 c.c.

108. 10 litres of gas are measured at 0° C. and 760 m.m. Bar.: what will they measure at 15° C. and 750 m m. Bar.?
Ans. 10·69 litres.

109. What is the difference between limestone, lime, and slaked lime? Give the formula of each.

110. By what process are stalactites formed?

SERIES XXIII.

111. A gas at 0° is raised in temperature to 12°, and then occupies the space of 1995 litres: what was its original volume? Ans. 1911 litres nearly.

112. A quantity of argentic oxide (OAg_2) on being heated yields 10 c.c. of oxygen: what was the weight of the oxide?
Ans. ·2078 gramme.

113. How much argentic nitrate (NO_2Ago) can be obtained by dissolving 1 oz. of pure metallic silver in nitric acid?
Ans. 1·574 oz.

114. What rate of crystallized baric chloride ($BaCL_2, 2OH_2$) can be obtained from 10 grammes of heavy spar (SO_2Bao'')? Ans. 10·47 grammes.

115. How much chalk will be formed by burning a gramme of charcoal and passing the product of combustion into an excess of lime water? Ans. $8 \cdot \dot{3}$ grammes.

SERIES XXIV.

118. If 1 cwt. of magnesian limestone ($C_2O_2Mgo''Cao''$) be treated with sulphuric acid, how much crystallized magnesic sulphate ($SOHo_2Mgo'',6OH_2$) will be obtained?
Ans. 149·73 lbs.

119. How much magnesia is contained in 10 grammes of magnesic pyrophosphate ($P_2O_3Mgo''_2$)?
Ans. 3·603 grammes.

120. The iron in 10 grammes of ironstone as a ferrous salt is found to reduce 2·81 grammes of potassic bichromate: what per-centage of iron does the ironstone contain? *
Ans. 32 per cent.

SERIES XXV.

121. What is the per-centage composition of the following substances?—

Calcic carbonate ($COCao''$). Ans. $\left\{\begin{array}{lr} Ca = & 40 \\ C = & 12 \\ O = & \underline{48} \\ & 100 \end{array}\right.$

Ammonic chloride (NH_4Cl). Ans. $\left\{\begin{array}{lr} N = & 26·17 \\ H = & 7·47 \\ Cl = & \underline{66·36} \\ & 100 \end{array}\right.$

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* $\left\{\begin{array}{l} CrO_2Ko \\ O \\ CrO_2Ko \end{array}\right. + 6FeCl_2 + 14HCl = 2KCl + Cr_2Cl_6 + 3FeCl_6 + 7OH_2.$

Example 16.—What is the per-centage composition of sodic carbonate ($CONao_2$)?

By the formula,—

$$\begin{array}{lr} C = & 12 \\ Na_2 = & 46 \\ O_3 = & \underline{48} \\ & 106 \end{array}$$

	$CONao_2$.		$CONao_2$.		C.		C.		Per centage composition.	
∴ As	106	:	100	::	12	:	x	=	11·32	C
					Na.		Na.			
	106	:	100	::	46	:	x	=	43·39	Na.
					O.		O.			
	106	:	100	::	48	:	x	=	45·28	O
									99·99	

Common salt ($NaCl$). Ans. $\left\{\begin{array}{l} Cl = 60{\cdot}68 \\ Na = \underline{39{\cdot}32} \\ \quad\quad 100 \end{array}\right.$

Sulphuric acid (SO_2Ho_2). Ans. $\left\{\begin{array}{l} S = 32{\cdot}65 \\ O = 65{\cdot}31 \\ H = \underline{\ \ 2{\cdot}04} \\ \quad\quad 100 \end{array}\right.$

Sodic nitrate (NO_2Nao). Ans. $\left\{\begin{array}{l} N = 16{\cdot}47 \\ O = 56{\cdot}47 \\ Na = \underline{27{\cdot}06} \\ \quad\quad 100 \end{array}\right.$

Cryolite ($6NaF, Al_2F_6$). Ans. $\left\{\begin{array}{l} Al = 13{\cdot}07 \\ Na = 32{\cdot}78 \\ F = \underline{54{\cdot}15} \\ \quad\quad 100 \end{array}\right.$

SERIES XXVI.

122. What are the empirical formulæ and names of the substances having the following per-centage composition?—

Example 17. A substance has been found to consist of calcium, phosphorus, and oxygen, in the following proportions:—

Calcium . .	38·71
Phosphorus . .	20·00
Oxygen . .	41·29
	100

a. $$\left.\begin{array}{lcr} \text{Carbon} & = & 42{\cdot}86 \\ \text{Oxygen} & = & \underline{57{\cdot}14} \\ & & 100 \end{array}\right\} \text{Ans. } CO.$$

b. $$\left.\begin{array}{lcr} \text{Silicon} & = & 20{\cdot}28 \\ \text{Magnesium} & = & 34{\cdot}16 \\ \text{Oxygen} & = & \underline{45{\cdot}56} \\ & & 100 \end{array}\right\} \text{Ans. } SiMgo'_2.$$

c. $$\left.\begin{array}{lcr} \text{Phosphorus} & = & 18{\cdot}90 \\ \text{Sodium} & = & 42{\cdot}07 \\ \text{Oxygen} & = & \underline{39{\cdot}03} \\ & & 100 \end{array}\right\} \text{Ans. } PONao_3.$$

d. $$\left.\begin{array}{lcr} \text{Iron} & = & 70{\cdot}01 \\ \text{Oxygen} & = & \underline{29{\cdot}99} \\ & & 100 \end{array}\right\} \text{Ans. } Fe_2O_3.$$

e. $$\left.\begin{array}{lcr} \text{Manganese} & = & 72{\cdot}05 \\ \text{Oxygen} & = & \underline{27{\cdot}95} \\ & & 100 \end{array}\right\} \text{Ans. } Mn_3O_4.$$

f. $$\left.\begin{array}{lcr} \text{Carbon} & = & 13{\cdot}04 \\ \text{Oxygen} & = & 52{\cdot}18 \\ \text{Magnesium} & = & 13{\cdot}04 \\ \text{Calcium} & = & \underline{21{\cdot}74} \\ & & 100 \end{array}\right\} \text{Ans. } C_2O_2Mgo''Cao''.$$

Now multiply by 2, to get rid of ·5, and we have 3 atoms of Ca, 2 atoms of P, and 8 atoms of O. The formula therefore is—

$$P_2O_8Ca_3 \text{ or } P_2O_2Cao''_3.$$

Example 18. What is formula of a substance having the following per-centage composition?—

$$\begin{array}{lcr} \text{Sodium} & = & 8{\cdot}73 \\ \text{Aluminium} & = & 10{\cdot}39 \\ \text{Silicon} & = & 32{\cdot}32 \\ \text{Oxygen} & = & \underline{48{\cdot}56} \\ & & 100 \end{array}$$

The relative number of atoms is obtained by dividing each per-centage amount by its atomic weight:—

$$\begin{array}{llllll} \text{Then } \dfrac{8{\cdot}73}{23} & = & {\cdot}379, & \text{relative number of atoms of} & & \text{Na.} \\ \dfrac{10{\cdot}39}{27{\cdot}5} & = & {\cdot}377, & ,, & ,, & \text{Al.} \\ \dfrac{32{\cdot}32}{28{\cdot}5} & = & 1{\cdot}134, & ,, & ,, & \text{Si.} \\ \dfrac{48{\cdot}56}{16} & = & 3.035, & ,, & ,, & \text{O.} \end{array}$$

These relative numbers are most readily reduced to their simplest relation by dividing each of them by the lowest, and allowing for errors of analysis, thus :—

$\frac{\cdot379}{\cdot377} = 1.$ No. of atoms of Na.

$\frac{\cdot377}{\cdot377} = 1.$,, ,, Al.

$\frac{1\cdot134}{\cdot377} = 3.$,, ,, Si.

$\frac{3\cdot035}{\cdot377} = 8.$,, ,, O.

The empirical formula of the substance is therefore Si_3NaAlO_8, and by doubling the formula in order that we may have the aluminic radical we have the constitutional formula $Si_6O_4Nao_2(Al_2O_6)^{vi}$.

FIRST STAGE OR ELEMENTARY EXAMINATION, MAY, 1871.

1. You have given to you some sulphur, water, and nitric acid. Describe how you would make sulphuric acid from these materials. (11.)
2. I mix together chlorine and hydrogen, and expose the mixture to sunlight; what happens? At the termination of the experiment I add ammonia to the product. What is the name and formula of the compound formed? (8.)
3. Express in words the meaning of the following equation :—

Hydrochloric acid
Water
Ammonia
Carbonic anhydride
Boric anhydride
Nitric acid
Ammonic nitrate
Sulphuretted hydrogen
Sulphurous acid
Sulphuric acid
Ammonic chloride. (12.)

8. How much, by weight, of each element is contained in 100 lbs. of a compound whose formula is $SO_4N_2H_8$? (9.)

9. How does boron occur in nature, and how would you prepare boric anhydride from borax? (8.)

10. If I burn a piece of sulphur in a closed flask filled with atmospheric air, what substances shall I find in the flask at the end of the operation, and what will be the relation between the volume of the gases originally contained in the flask and that of the gases found there at the close of the experiment? If I repeat the experiment in a flask filled with pure oxygen instead of atmospheric air, what difference shall I find in the composition and volume of the resulting gases? (15.)

11. Classify twenty-two of the most important elements into metals and non-metals; also into positive and negative elements. (12.)

12. You have some ammonic carbonate (carbonate of ammonia) and nitric acid, and are required to make and collect laughing gas from these materials. How would you do it? Describe minutely the apparatus you would employ, and make a sketch of it. What are the chief properties of laughing-gas? (14.)

SECOND STAGE OR ADVANCED EXAMINATION.

20. In dissolving zinc in hydrochloric acid, and collecting the hydrogen evolved, 250 cubic centimetres of the gas measured at 0° C. and 760 m.m. pressure were obtained? How much zinc was dissolved? (9.)

21. I pass sulphuretted hydrogen through an aqueous and acid solution of the following salts :—

Ferric chloride.
Chrome alum.
Cadmic sulphate.
Potassic chromate.
Stannic chloride.
Magnesic sulphate.

Explain the changes which occur, and draw graphic formulæ of the products. (15.)

22. You are required to perform a qualitative analysis of a solid substance containing the following ingredients :—Baric

carbonate, sodic sulphate, cupric chloride, zincic chloride, and magnesic chloride. State exactly how you will separate and individually recognise each constituent. (16.)

23. What substances give to water the quality of hardness, and how do you distinguish between temporary hardness and permanent hardness? (11.)

24. What do you mean by absolute, latent, and active atomicity? Illustrate your meaning by examples. (11.)

25. You have some nitric acid and metallic copper given to you, and are required to prepare nitrogen gas from these materials. How will you do it? (12.)

26. I heat each of the following substances to redness in a porcelain crucible; what takes place? Illustrate all chemical changes by equations:—Ammonic nitrite, ammonic nitrate, hydric disodic phosphate, sodic chloride, mercuric iodide, platinic chloride, plumbic nitrate, argentic oxide. (13.)

27. Describe the process of manufacturing soda ash, as carried on in the alkali works of this country. (13.)

28. How would you prepare a specimen of pure silica from felspar? (10.)

29. How would you prepare hydriodic acid, and fill a glass cylinder with it? (9.)

APPENDIX.

Weights.

1 gramme = the unit (15·434 grains).
10 grammes = 1 decagramme.
100 ,, = 1 hectogramme.
1000 ,, = 1 kilogramme.
·1 gramme = 1 decigramme.
·01 ,, = 1 centigramme.
·001 ,, = 1 milligramme.

Linear Measure.

1 metre = the unit (39·37 inches).
10 metres = 1 decametre.
100 ,, = 1 hectometre.
1000 ,, = 1 kilometre.
·1 metre = 1 decimetre.
·01 ,, = 1 centimetre.
·001 ,, = 1 millimetre.

Measure of Capacity.

1 litre = the unit (61·027 c.in., or 1·76 pints).
10 litres = 1 decalitre.
100 ,, = 1 hectolitre.
1000 ,, = 1 kilolitre (1 cubic metre).
·1 litre = 1 decilitre.
·01 ,, = 1 centilitre.
·001 ,, = 1 millilitre (1 cubic centimetre).

1 cubic centimetre of distilled water, taken at 4° C., its point of greatest density, weighs 1 gramme.
1 litre of distilled water, taken at 4° C., weighs 1000 grammes.
1 litre of hydrogen at 0° C. and 760 m.m. Bar. weighs 1 crith.

1 Crith = ·0896 Gramme.

SYLLABUS OF THE SCIENCE AND ART DEPARTMENT.

Inorganic Chemistry.

FIRST STAGE OR ELEMENTARY COURSE.

Pupils presenting themselves for examination will be expected to possess a knowledge of the following subjects :—

Definition of chemistry. Simple and compound matter. Different modes of chemical action. Combining weights. Volume Weights. Principles of chemical nomenclature. Symbolic notation. Graphic notation. Chemical formulæ. Chemical equations. Atomicity of elements. Simple and compound radicals. Definition of a compound radical. Classification of elements into metals and non-metals, into chlorous and basilous elements. Classification according to atomicity.

French and English systems of weights and measures. Conversion of English into French weights and measures. The crith and its uses.

Hydrogen.—Its preparation and properties.

Chlorine.—Preparation of chlorine from hydrochloric acid. Analysis and synthesis of hydrochloric acid. Properties and reactions of hydrochloric acid.

Oxygen.—Its preparation and properties. Allotropic oxygen or ozone. Formation and reactions of water. Preparation and properties of hydroxyl. Oxides and oxy-acids of chlorine.

Boron.—How it occurs in nature. Its allotropic modifica-

Expansion of gases by heat. Reduction of gaseous volumes to standard pressure and temperature.

Manufacture of hydrochloric, nitric, and sulphuric acids. Composition and manufacture of bleaching powder. Theory of bleaching. Suitability of water for domestic purposes. Causes of permanent and temporary hardness in water.

Bromine.—Hydrobromic and bromic acid.

Iodine.—Hydriodic, iodic, and periodic acid.

Fluorine.—Hydrofluoric acid.

Silicon.—Silica. Silicic acid. Silicic hydride. Names and formulæ of some of the more important siliceous minerals.

Phosphorus.—Phosphoretted hydrogen. Acids and anhydrides of phosphorus.

Arsenic.—Arsenious and arsenic acids. Arseniuretted hydrogen. Detection of arsenic.

Antimony and Bismuth.—Preparation and properties of their chief compounds.

The monad metals, especially potassium, sodium, and silver. Manufacture of soda ash.

The diad metals—barium, strontium, calcium, magnesium, zinc, cadmium, mercury, and copper.

The chief properties of the following metals:—gold, aluminium, platinum, lead, chromium, manganese, iron, cobalt, and nickel.

Composition, preparation, and properties of the more important compounds of these metals.

Outline of qualitative analysis. Reaction of the principal mineral acids and bases. Course pursued in the application of these reactions to the analysis of a mixture of several acids and bases.

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,, ,, com., per gal.	0 0	6 0
,, Citric	0 4	3 6
,, Hydrochloric	0 0	0 2
,, ,, pure	0 1	0 8
,, Hydrofluoric	0 2	1 6
,, Hydrofluosilicic	0 2	2 0
,, Hydrosulphuric (Sulphuretted Hydrogen Solution)	0 1	0 6
,, Nitric	0 0	0 4
,, ,, pure, sp. gr. 1·42	0 1	1 0
,, ,, ,, ,, 1·50	0 3	2 6
,, Picric (Carbazotic)	0 6	5 0
,, Pyrogallic	4 0	0 0
,, Sulphuric	0 0	0 2
,, ,, pure	0 1	1 0
,, ,, Nordhausen	0 2	1 6
,, Sulphurous	0 2	1 0
,, Tannic	1 0	0 0
,, Tartaric (Howard's)	0 3	2 6
Alum	0 0	0 3
,, Purified	0 1	0 8
,, Chrome	0 2	1 6
,, Iron, Ammonia	0 2	1 6
Aluminium, Sulphate, purified	0 0	1 6
Ammonium, Carbonate	0 1	1 0
,, Chloride	0 1	0 8
,, ,, pure	0 2	1 4
,, Hydrosulph.	0 2	1 4
,, Molybdate	2 6	0 0
,, Nitrate	0 2	2 0
,, Oxalate	0 4	4 0
,, Sulphocyanide	0 4	4 0
Ammonia, liquid, sp. gr. ·880	0 1	1 0
Antimony, Metal	0 2	1 6
,, Sulphide, Powder	0 1	0 8
Asbestos	0 3	3 0
Barium, Chloride	0 1	1 0
,, ,, pure	0 2	1 6
,, Carbonate	0 2	2 0
,, Nitrate	0 2	1 6
,, *Oxide*	0 6	6 0
Barium, Peroxide	1 0	0 0
,, Sulphate	0 2	2 0
Bismuth, Trisnitrate	2 0	0 0
Borax	0 1	1 0
Bromine	1 6	0 0
Cadmium, Iodide	2 0	0 0
Calcium, Carbonate, White Marble	0 0	0 4
,, Nearly white do.	0 0	0 2
,, Chloride, fused white	0 2	1 6
,, Chloro-hypochlorite ,, Bleaching Powder	0 0	0 4
Carbon, Disulphide	0 1	1 0
Charcoal, Animal	0 0	0 3
,, ,, grain	0 0	0 6
,, Wood, for blowpipe	0 0	0 3
Chlorine, Solution	0 1	1 0
Chromium, Oxide, Green	0 6	0 0
,, Sulphate	1 0	0 0
Cobalt, Nitrate	2 0	0 0
,, ,, Solution	0 6	0 0
,, Oxide	2 0	0 0
Copper, Carbonate	0 6	0 0
,, Chloride	0 4	0 0
,, Nitrate	0 3	2 6
,, Oxide, Black	0 3	2 6
,, ,, Red	0 4	4 0
,, Sulphate	0 0	0 5
,, ,, pure	0 2	1 6
,, Turnings	0 1	1 4
Fluor Spar	0 0	0 4
Glycerine	0 3	0 0
Indigo, Sulphate, Solution	0 1	1 0
Iodine	1 4	0 0
,, Resublimed	2 0	0 0
Iron, Alum	0 2	1 6
,, Ammonio-sulphate	0 2	1 0
,, Borings	0 0	0 3
,, Perchloride	0 3	2 6
,, Sesquioxide	0 1	1 0
,, Sulphate (Copperas)	0 0	0 2
,, ,, pure	0 0	0 6
,, Sulphide	0 1	0 6
Lead, Acetate	0 1	0 [illegible]

	Per Oz. s. d.	Per Lb. s. d.
Lead, Acetate, pure	0 2	1 6
,, Carbonate, precipitated	0 2	2 0
,, Nitrate	0 1	1 0
,, ,, pure	0 2	2 0
,, Oxide (Litharge)	0 0	0 6
,, Red Oxide (Red Lead)	0 0	0 4
,, Tartrate	0 4	0 0
Lime, Caustic (Derbyshire)	0 0	0 2
,, Bisulphate, Solution	0 0	0 6
Litmus	0 4	0 0
,, Books each	0 2	0 0
,, Tincture	0 4	0 0
,, Infusion	0 2	0 0
Magnesium Ribbon.. per yd.	0 5	0 0
,, Carbonate	0 1	1 0
,, ,, heavy	0 2	2 0
,, Sulphate	0 0	0 6
Manganese, Carbonate	0 4	4 0
,, Chloride	0 4	4 0
,, Sulphate	0 6	0 0
,, Peroxide, Powder	0 0	0 3
Microcosmic Salt	0 6	0 0
Nessler Test	0 6	0 0
Nickel, Sulphate	1 0	0 0
,, Metallic	1 0	0 0
Phosphorus	0 4	0 0
,, Amorphous (com.)	1 0	0 0
Platinum Bichloride .. dram	4 0	0 0
Potassium ....2s. per dram	12 0	0 0
,, Hydrate (Caustic Potash)	0 4	4 0
,, Acetate	0 3	2 6
,, Bicarbonate, Crystals	0 2	1 6
,, Bichromate	0 0	0 6

	Per Oz. s d.	Per Lb. s. d.
Sodium, Caustic (Hydrate)	0 0	0 6
,, Hyphosulphite	0 0	0 6
,, Nitrate, pure	0 2	1 6
,, Phosphate	0 2	1 6
,, Sulphate	0 0	0 4
,, Bisulphite Solution	0 1	1 0
Strontium, Chloride	0 2	1 6
,, Nitrate	0 1	1 0
Sulphur, Roll	0 0	0 3
Sulphuretted Hydrogen Solution	0 1	0 6
Tin, Grain	0 2	1 8
,, Granulated	0 2	2 0
,, Chloride Solution	0 1	0 8
Zinc, in Narrow Clippings	0 0	0 2
,, free from Arsenic	0 3	3 0
,, Granulated	0 0	0 4
,, Carbonate	0 4	0 0
,, Sulphate, pure	0 2	1 0
,, Oxide	0 3	0 0

SAMPLES OF DYE-WARES FOR EXPERIMENTAL DYEING.

Annatto	2 0	2 0
Aniline Colours—Red (Magenta)	1 0	0 0
Violet	2 0	0 0
Blue	2 0	0 0
Green	2 0	0 0
Bismarck Brown	1 0	0 0
Barwood	0 0	0 2
Camwood	0 0	0 6
Catechu	0 0	0 4
Cochineal	0 4	4 0